By Matthew Raspanti

Chapter 15

THE QUANTUM LEAP

"*A new scientific truth does not triumph by convincing its opponents and making them see the light, but rather because its opponents eventually die, and a new generation grows up that is familiar with it.*"

Max Planck, German physicist (1858-1947)

SPECTRAL LINES

Around 1666, you will recall, Newton discovered that, after a narrow beam of sunlight passes through a glass prism, a series of colored bands appears on a screen in the familiar sequence of the rainbow; he called it the "spectrum" of sunlight. In 1802, another Englishman, W.H. Wollaston, observed that there were some dark lines among the intense colors of this spectrum.

In 1814, a German physicist, Joseph von Fraunhofer, modified Newton's experiment by replacing the screen with a telescope through which he could view sunlight that had passed through a narrow slit and a glass prism. Through his "spectrometer", he could see a multicolored strip displaying here and there "an almost countless number of strong or weak lines which are darker than the rest of the colored image; some appeared to be almost perfectly black."

An explanation for the Fraunhofer lines was found by another German physicist, Gustav Robert Kirchhoff. In 1859, he did an experiment in which light from the Sun was made to pass through a flame heavily charged with salt (sodium chloride), before going through the narrow slit of a spectrometer. Within the solar spectrum, two bright lines appeared in place of the two closely spaced dark lines that Fraunhofer had labeled the "D lines".

Kirchhoff concluded that sodium present in the Sun's atmosphere absorbs the radiation corresponding to the two D lines. In his experiment, the sodium in the flame had emitted radiation that restored the same two lines. The other Fraunhofer dark lines could be similarly attributed to the absorption by other elements present in the Sun's atmosphere.

Shortly after, Kirchhoff announced the two fundamental laws of the new science of "spectroscopy":

- Each chemical element has a unique spectrum, which consists of a number of distinctive lines. **Figure 15.1** shows, for instance, the spectral lines that are unique to sodium, hydrogen and mercury. Each line corresponds to a particular frequency of electromagnetic radiation.
- Each element is capable of absorbing the same radiation it is capable of emitting.

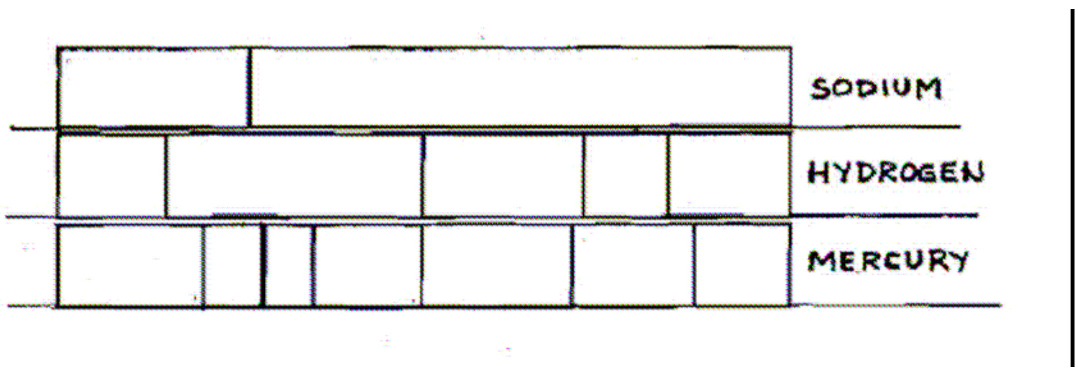

Figure 15.1 - Spectral lines of three elements
(From Cooper Plate 34.3 pg 433)

It was soon discovered that some of the lines unique to an element were in the infrared or ultraviolet, rather than the visible, portion of the spectrum.

At the time, physics had no explanation for the amazing fact that each element has a unique set of fingerprints. The answer came half a century later from quantum theory.

PLANCK'S QUANTUM

As we heat an object, say, the tip of an iron rod in a blacksmith's forge, the light it emits turns from dull red to bright red and then to white. If heated to a sufficiently high temperature, it turns bluish. What causes this gradual change of colors?

In the late 1800's, physicists studied extensively how electromagnetic radiation, including light, is emitted by a heated oven through a small hole. Since the molecules of the heated material inside the oven are vibrating, charged particles within the molecules are in accelerated motion, causing the emission of electromagnetic waves. These are repeatedly bounced around by the walls of the oven, until they eventually escape through the small hole As the oven is heated to higher temperatures, the hole will glow first dull red, then bright red, then white, then bluish, which is also what happens to the heated tip of an iron rod.

The radiation escaping from the hole can be analyzed experimentally to determine how the brightness of the light changes at various frequencies for a given temperature of the oven. Plotting the data results in curves like the two illustrated in **Figure 15.2**.

Attempts were made to predict mathematically the shape of these curves. Theoretical studies, however, concluded that, at any temperature, the intensity of the radiation emitted should continue to increase with the frequency, instead of decreasing after reaching a peak. Most of the radiation should be in the ultraviolet region of the spectrum and beyond. This result, first obtained in 1900, was clearly contrary to the experimental evidence. It became known as the "ultraviolet catastrophe".

In Germany, Max Planck (1858-1947) was able to derive a mathematical equation that was in agreement with the experimental curves. To succeed, however, he had to make a peculiar assumption, just tailored to get the desired result.

At the time, energy was thought to be continuous, i.e., it could increase or decrease by infinitesimal amounts. In terms of currency, we might compare it to very fine gold powder, with which we can pay in whatever amounts may be needed [5]. To get the desired result, however, Planck had to assume that energy is discontinuous, more like a currency with coins and bills of only certain denominations. To pay for some transactions, then, we have to use whole multiples of a single coin or bill of the right denomination.

Planck had to assume that heat converts into light not in any arbitrarily small increments, but in a granular fashion. The smallest quantity, or "quantum", of heat energy that can be converted into light of frequency f had to be equal to the frequency f times some constant h, h x f. (The constant h, called Planck's constant, would become as important in physics as the constant c, the speed of light in a vacuum.)

The total amount of heat energy converted into light of some frequency is a whole multiple of individual "granules" or quanta whose size is proportional to the frequency.

[5]. For gold powder to be truly continuous, its grains would have to be infinitesimally small.

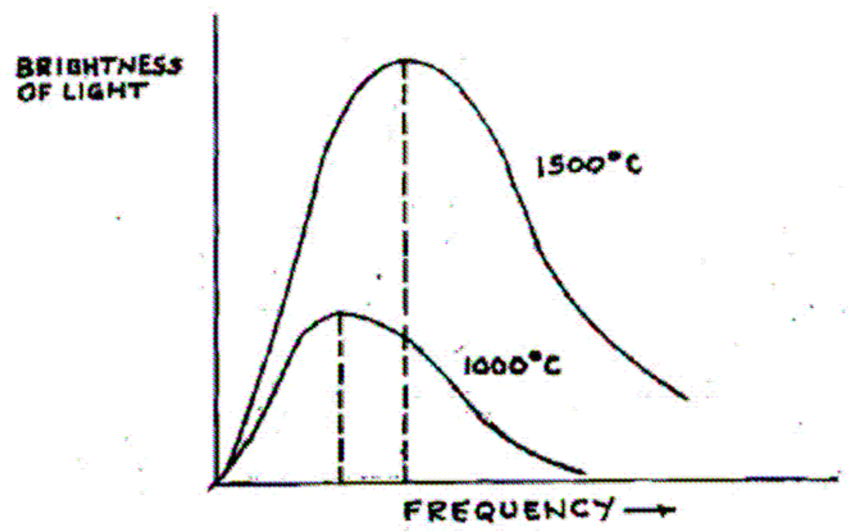

Figure 15.2 - Electromagnetic radiation emitted at two different temperatures

In the lower curve, the brightness of the emitted light first increases with frequency toward a peak, and then starts decreasing.

The upper curve describes what happens for a higher temperature of the oven. The range of frequencies becomes wider, the values of brightness are higher, and so is the frequency at which maximum brightness occurs. As the temperature rises, the curve shifts toward the right (in the direction from red to blue to violet). (from Poet pg 189)

We can see now, roughly, how Planck's assumption can explain the shape of the previous plots. After reaching a peak, each curve starts to decline because of the increasing investment of energy required to create a quantum at higher frequencies. Neither small coins nor very large bills are likely to contribute to the bulk of the money to be found in a cash register.

The assumption of granules or "quanta" of energy gave the desired result, but appeared to be totally unnatural to the physicists of the time, Planck included. In classical theory, the energy of a wave depended on the amplitude of the wave, not on its frequency. For years, the idea of the quantum was viewed - even by Planck - as a convenient "working hypothesis", a mathematical trick with no physical basis.

Planck, who died at 89 in 1947, endured many personal tragedies. Both of his daughters died in childbirth shortly after getting married. His elder son was killed in battle during World War I. His other son was implicated in an attempt to assassinate Hitler, and died a horrible death at the hands of the Gestapo. His house was destroyed by an air raid.

EINSTEIN'S PHOTON

The next application of Planck's quantum hypothesis was made by Einstein. He used it to explain a relatively obscure phenomenon called the "photoelectric effect".

In the late 1800's, it was known that, under certain conditions, when light was shined on a metal surface, it would cause electrons to be knocked off the surface. Some qualitative facts were known, but without any explanation:

- The photoelectric effect is easily produced with blue or ultraviolet light, but not with red light, which has a lower frequency.
- When electrons are emitted from the metal surface, their kinetic energy does not exceed some maximum value. If we increase the *brightness* of the light, more electrons are emitted, but their maximum kinetic energy does not increase. It will increase, however, if we increase the *frequency* of the light.

Einstein gave an explanation based on Planck's hypothesis, which he viewed, not as a mathematical trick, but as a profound truth applicable to all aspects of physics.

In one of his famous papers of 1905, he suggested that light was not only emitted in Planck's little granules of energy, but it was also absorbed in such granules. This granularity represented a property of light itself, independent of how it was emitted or absorbed. Einstein proposed a theory in which light can be likened to a hail of energy granules with no mass. These energy granules were later named "photons", each carrying a quantum of energy proportional to the frequency of the light.

The theory provided a convincing explanation of the photoelectric effect. A single electron will either absorb a whole photon or none at all. Only if the frequency of the light is high enough, will the "kick" delivered by the photon be large enough to enable the stricken electron to break free from the metal, after colliding with atoms, and have some kinetic energy left over.

If the frequency of the light is even higher, the photon's "kick" will be stronger and the stricken electron will have more kinetic energy left over, after breaking free. Increasing the intensity of the light, instead of its frequency, makes more photons available, but does not increase the strength of the "kick" each photon can deliver.

The puzzle of the photoelectric effect had been solved, but only to raise new questions. How could Einstein's particle theory of light be reconciled with the overwhelming evidence that supported the wave theory of light? How could light be both waves *and* particles? This was the first instance of the paradoxical wave-particle duality that would become a fundamental aspect of 20th-century physics.

When Einstein was submitted in 1913 for membership in the Prussian Academy of Sciences, his sponsors felt obliged to make excuses for what seemed then to be the least respectable of his many achievements, the photoelectric theory. Three years later, however, the American physicist Robert Millikan experimentally confirmed Einstein's quantitative predictions. (Ironically, he had started with the intent of proving them wrong.)

In 1921, Einstein received the Nobel Prize in physics for his "photoelectric law and his work in the field of theoretical physics". No mention was made of relativity, still a controversial subject.

BOHR'S ATOM

After Rutherford proposed an atom with a central nucleus, a natural next step was to imagine the atom as a miniature solar system. Instead of planets orbiting the Sun under the pull of gravity, one would have negative electrons orbiting the positive nucleus under the pull of electrical attraction.

The idea of such a parallel between the very small and the very large was very appealing, but totally unworkable. According to Maxwell's theory, since an electron would be continuously accelerated as it revolved around the nucleus, it should radiate light and thereby lose energy. This would cause it to spiral faster and faster toward the center; within billionths of a second, it would crash into the nucleus.

A new model of the atom was proposed in 1913 by a young Danish physicist, Niels Bohr (1885-1962). After receiving his doctorate at the University of Copenhagen, in 1912 Bohr went to work under Rutherford in Manchester. Here, within three months, he laid the foundations for a new theory of matter that would break away completely from what is now called classical physics.

Bohr, who was familiar with Planck's work, started with the simplest of all atoms, the hydrogen atom, which consists of a single proton and a single electron. To stay consistent with experimental facts, he combined some well-established principles of classical physics with some nonclassical hypotheses, which, like Planck's, were purposely tailored to fit the situation. The theory that resulted was a hodgepodge, but it worked.

For his model of the hydrogen atom, Bohr made three assumptions:

- Electrons move in circular orbits, but only orbits with certain radii are allowed.
- As long as an electron is moving in an allowed orbit, it does not radiate energy, as it should according to Maxwell's theory.
- Energy is radiated only when an electron leaps from one allowed orbit to another.

For the smallest allowed orbit of the electron, orbit #1, Bohr computed its radius to be about 5 billionths of a centimeter, about 50,000 times the radius of the nucleus. When the electron is in this orbit, it is said to be in its "ground state". It is as close as it can ever get to the nucleus.

For orbit #2, the radius is 4 (2x2) times the minimum radius; for orbit #3, it is 9 (3x3) times; for orbit #4, it is 16 (4x4) times, and so on.

In Bohr's theory, there is an "energy level" associated with each allowed "state" or orbit of the electron. You can think of this energy level as what the electron "owes" to become free from the nucleus. It is considered to be negative energy in the same sense that debt can be viewed as negative wealth: the smaller your debt, the larger your "wealth".

When the electron is in orbit #1, it is at its lowest energy level (highest debt). In orbit #2, its debt is 4 (2x2) times smaller; in orbit #3, it is 9 (3x3) times smaller, and so on. The larger the radius of the orbit, the more loosely bound the electron is.

The electron is normally in its smallest orbit at its lowest energy level. Let us assume that, while the electron is in this state, the hydrogen atom is disturbed somehow, for instance, by collisions with other atoms, or by photons bombarding it. If the electron receives a shot of energy equal to the difference between energy levels #1 and #2, it will be "excited" to jump to level #2.

Conversely, when the electron makes a transition from level #2 back to level #1, it releases a photon whose energy is the difference of the two energy levels (and whose frequency is its energy divided by Planck's constant).

More generally, if the electron receives a shot of energy in just the right amount, it can make a "quantum leap" from whatever energy level it already occupies to a higher one, not necessarily consecutive. Later, it can slide down to some lower level, not necessarily consecutive, and emit a photon, whose energy is the difference of the two levels.

We might say that, in making its transitions up or down, the electron, instead of a continuous ramp, has available a ladder with discrete rungs at uneven distances. When it leaps from one rung, it must land on another.

In Bohr's theory, consecutive orbits of the electron are assigned consecutive integer values of what is called the "quantum number n". That is why we have spoken of orbits #1, #2, #3 and so on. For each value of the "quantum number n", Bohr was able to compute the values of the associated radius and energy level. **Figure 15.3** shows the "rungs" of the energy "ladder" of the hydrogen atom for values of the quantum number n from 1 to 5.

Bohr's theory could explain with remarkable accuracy many details of the hydrogen spectrum, including spectral lines that were discovered outside the visible spectrum.

In 1913, Bohr published his theory of the hydrogen atom. In 1916, at 31, he returned to Denmark to accept a professorship at the University of Copenhagen. In 1920, he became director of the newly founded Institute of Theoretical Physics, which over the years attracted many of the world's leading physicists to Copenhagen. In 1922, he was awarded the Nobel Prize.

When Denmark was invaded by Germany in 1940, Bohr refused to cooperate with the Nazis. In 1943, under threat of imminent arrest, he was whisked away with his family to Sweden on a fishing boat in the middle of the night by the Danish underground. Eventually, he made his way to the United States, where he worked at the Los Alamos Laboratory on the Manhattan Project, which built the first atom bomb. He later returned to Copenhagen, where he died in 1962.

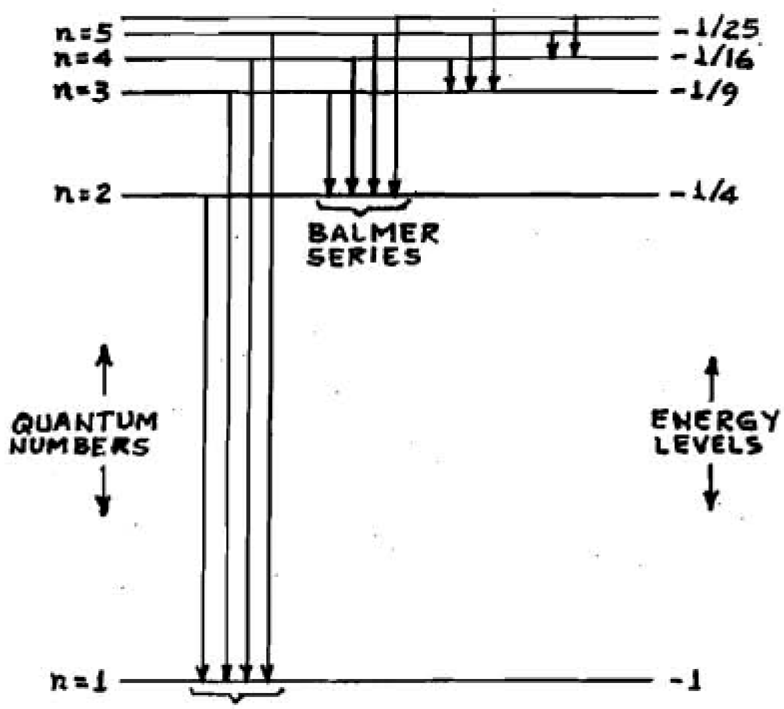

Figure 15.3 - Energy level diagram for hydrogen in Bohr's model

For the various values of the quantum number n, the corresponding energy levels are shown relative to the lowest level, to which is accordingly assigned the value -1.

The arrows represent downward transitions between energy levels. Each corresponds to a line in the hydrogen spectrum. The four transitions down to level #2 correspond to the four spectral lines shown in Figure 15.1. These are the only ones to fall within the visible spectrum.

(From Hey Figure 4.13 pg 48)