Lecture 3: Electrons in Atoms

Discovery of electron

In 1897, when the first experimental evidence for the internal structure of atoms was discovered when the British physicist J. J. Thomson was investigating “cathode rays,” the rays that are emitted when a high potential difference (a high voltage) is applied between two metal electrodes in an evacuated glass tube. Thomson showed that cathode rays are streams of negatively charged particles coming from inside the atoms that made up the negatively charged electrode, the cathode. Thomson found that the charged particles, which came to be called electrons, were the same regardless of the metal he used for the cathode. He concluded that they are part of the makeup of all atoms. Thomson was able to measure the value of e/me, the ratio of the magnitude of the electron’s charge e to its mass m. Mulliken later determined charge (oil drop experiment).

Rutherford experiment

Although electrons have a negative charge, an atom has zero charge. Therefore, an atom must contain enough positive charge to cancel the negative charge. But where was the positive charge? Thomson suggested a model of an atom as a blob of a positively charged, jellylike material, with the electrons suspended in it like raisins in pudding. However, this model was overthrown in 1908 by another experimental observation. Ernest Rutherford knew that some elements, including radon, emit streams of positively charged particles, which he called particles (alpha particles). He asked two of his students, Hans Geiger and Ernest Marsden, to shoot particles toward a piece of platinum foil only a few atoms thick. If atoms were indeed like blobs of positively charged jelly, then all the particles would easily pass through the diffuse positive charge of the foil, with only occasional slight deflections in their paths. Geiger and Marsden’s observations astonished everyone. Although almost all the particles did pass through and were deflected only very slightly, about 1 in 20 000 was deflected through more than 90º, and a few particles bounced straight back in the direction from which they had come. “It was almost as incredible,” said Rutherford, “as if you had fired a 15-inch shell at a piece of tissue paper and it had come back and hit you.”

Planetary model fails

It failed since accelerating electron in a curved orbit will continuously emit radiation- classical EM theory. Hence the electron would immediately collapse into the nucleus.

Photoelectric effect

Electrons emitted from a metal surface upon irradiation above a certain frequency. The kinetic energy increases with frequency above the threshold frequency. This is contrary to what is known in classical physics.

Classical physics predicted that kinetic energy would not change with the frequency of light. Nothing in classical physics connected frequency to energy. In addition, it predicted that KE should be dependent on the Intensity of light. The observation was the KE of electrons had nothing to do with the intensity of light. It was a real conundrum.

The KE vs frequency plot for different metals were same and found to be equal to Planks constant.

Just a few years earlier, Max Plank was interested in the origin of black body radiation.

Black Body Radiation

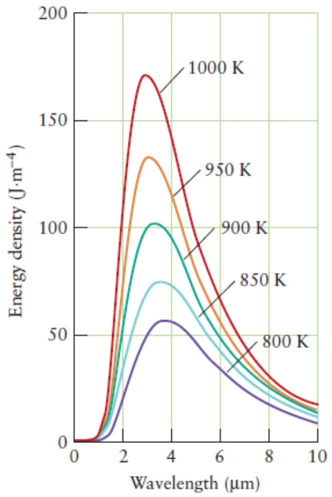


FIGURE 1.12 The intensity of radiation emitted by a heated black body as a function of wavelength. As the temperature increases, the total energy emitted (the area under the curve) increases sharply, and the maximum intensity of emission moves to shorter wavelengths.

For nineteenth-century scientists, the obvious way to account for the laws of black-body radiation was to use classical physics, the theory of motion devised by Newton two centuries previously, to derive its characteristics. However, much to their dismay, they found that the characteristics they deduced did not match their observations. Worst of all was the ultraviolet catastrophe: classical physics predicted that any hot body should emit intense ultraviolet radiation and even x-rays and γ-rays! According to classical physics, a hot object would devastate the countryside with high-frequency radiation. Even a human body at 37 C would glow in the dark. There would, in fact, be no darkness.

Plank's hypothesis

Plank proposed that in these materials there must be oscillators which must be giving of radiation and this radiation was being emitted in quanta or chunks. Using this idea and some statistical mechanics he was able to calculate the shapes of these curves. He got the shape correct but to get the intensity correct he had to use scaling factors in front of the frequency of his oscillators and that is what we know as the Plank's constant.

Einstien's hypothesis of the PE effect

Einstien was amazed that the same number comes in the slope of graph of KE vs phtoelectric effect. He then put this into the equation: $E = h\nu - h\nu$. Since ν is the energy of the incident radiation, $h\nu$ is the the energy of the incident radiation. That is how $E=h\nu$ comes up- from photoelectric effect.

In classical EM there is no relation between Energy and frequency. You can have any frequency but the energy comes quantized because of the factor h. Later it was called the energy of photon.

For PE effect, you have to get a packet of energy or a quantum of energy i.e. sort of a particle. The first evidence of particle nature of light. Eisntein went on to show a more convincing example for the particle like nature nature of radiation….he showed that a particle has momentum. Even though a photon doesn't have mass- rest mass as the physicists would say. Momentum $p = mv$. What Einstein showed from the relativistic equation of motions, that a photon having frequency γ has a momentum

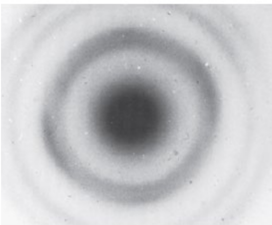
$p=h\nu/c$. Since $\lambda = c/\nu$, we can write: $p = h/\lambda$. That is the photon momentum. The Compton experiment demonstrated that photon had momentum.

How does the light behave during experimental conditions- as a wave or as a particle? There is no contradiction- it is the fundamental property of radiation.

Lecture 4: Matter as waves

Davison and Grammer experiment

The wavelike character of electrons was confirmed by showing that they could be diffracted. The experiment was first performed in 1925 by two American scientists, Clinton Davisson and Lester Germer, who directed a beam of fast electrons at a single crystal of nickel. The regular array of atoms in the crystal, with centers separated by 250 pm, acts as a grid that diffracts waves; and a diffraction pattern was observed. Since then, heavier particles, such as molecules, have also been shown to undergo diffraction, and there is no doubt that particles have a wavelike character. Indeed, electron diffraction is now an important technique for determining the structures of molecules and exploring the structures of solid surfaces.



What about your and mine wave-length. The wavelength of cricket ball – it is 10^{-34} m. The size of a nucleus- 10^{-14}m. Hence this will not have any consequence in our macroscopic world. Because to see any effects from this small wave-length, we need to have two nickel atoms or slits as small as these.

Uncertainty Principle

The discovery of wave–particle duality not only changed our understanding of electromagnetic radiation and matter, it also swept away the foundations of classical physics. In classical mechanics, a particle has a definite trajectory, or path on which location and linear momentum are specified at each instant. Think of the trajectory of a ball: in principle, we can state its location and momentum at every moment of its flight. However, we cannot specify the precise location of a particle if it behaves like a wave: think of a wave in a guitar string, which is spread out all along the string, not localized at a precise point. A particle with a precise linear momentum has a precise wavelength; but, because it is meaningless to speak of the location of a wave, it follows that we cannot specify the location of a particle that has a precise linear momentum. The wave–particle duality of matter means that the electron in a hydrogen atom cannot be described as orbiting the nucleus with a definite trajectory.
The popular picture of an electron in orbit around the nucleus is just plain wrong! The difficulty will not go away. Wave–particle duality denies the possibility of specifying the location if the linear momentum is known, and so we cannot specify the trajectory of any particles exactly. The uncertainty is negligible for heavy particles, but for subatomic particles it can be huge. Thus, if we know that a subatomic particle is *here* at one instant, we can say nothing about where it will be an instant later! The impossibility of knowing the precise position if the linear momentum is known precisely is an aspect of the complementarity of location and momentum— if one property is known the other cannot be known simultaneously. The Heisenberg uncertainty principle, which was formulated by the German scientist Werner Heisenberg in 1927, expresses this complementarity quantitatively. It states that, if the location of a particle is known to within an uncertainty *x*, then the linear momentum, *p*, parallel to the *x*-axis can be known simultaneously only to within
an uncertainty *p*, where

$$\overbrace{\Delta p}^{\text{Uncertainty in momentum}} \times \overbrace{\Delta x}^{\text{uncertainty in position}} \geq \tfrac{1}{2}\hbar \qquad (8)^*$$

The symbol $\hbar$, which is read "h bar," means $h/2\pi$, a useful combination that is found widely in quantum mechanics.

The uncertainty principle has negligible practical consequences for macroscopic objects, but it is of profound importance for very precise measurements dealing with subatomic particles, such as the locations and momenta of electrons in atoms, and the interpretation of their properties.