Every so often you hear a commercial or a news story with the words "quantum leap" in it. The quantum leap is supposed to be a major breakthrough, a big change, something extraordinarily large. The reality is far different. Instead of the big, extravagant change the "quantum" that scientists know about is a very small difference in the location of an electron around a nucleus - hardly an enormous shift at all.

Quantization of Energy

German physicist Max Planck (1858 - 1947) studied the emission of light by hot objects. You have likely see a heated metal object glow and orange-red color (see below).

![Heated object glowing](image)

A heated object may glow different colors. The atoms in this piece of metal are releasing energy in discrete units called quanta.

Classical physics, which explains the behavior of large, everyday objects, predicted that a hot object would emit electromagnetic energy in a continuous fashion. In other words, every wavelength of light could possibly be emitted. Instead, what Planck found by analyzing the spectra was that the energy of the hot body could only be lost in small discrete units. A quantum is the minimum quantity of energy that can either be lost or gained by an atom. An analogy is that a brick wall can only undergo a change in height by units of one or more bricks and not by any possible height. Planck showed that the amount of radiant energy absorbed or emitted by an object is directly proportional to the frequency of the radiation.

\[ E = h \nu \]

In the equation, \( E \) is the energy, in joules, of a quantum of radiation, \( \nu \) is the frequency, and \( h \) is a fundamental constant called Planck's constant. The value of Planck's constant is \( h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \). The energy of any system must increase or decrease in units of \( h \nu \). A small energy change results in the emission or absorption of low-frequency radiation, while a large energy change results in the emission or absorption of high-frequency radiation.
What is the energy of a photon of green light with a frequency of \(5.75 \times 10^{14} \text{ Hz}\)?

**Solution:**

**Step 1: List the known quantities and plan the problem.**

**Known**
- Frequency \(\nu = 5.75 \times 10^{14} \text{ Hz}\)
- Planck's constant \(h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}\)

**Unknown**
- Energy \(E\)

Apply the equation \(E = h \nu\) to solve for the energy.

**Step 2: Calculate**

\[
E = (6.626 \times 10^{-34} \text{ J} \cdot \text{s}) \times (5.75 \times 10^{14} \text{ Hz}) = 3.81 \times 10^{-19} \text{ J}
\]

**Step 3: Think about your result.**

While the resulting energy may seem very small, this is for only one photon of light. Visible quantities of light consist of huge quantities of photons. Recall that a hertz is equal to a reciprocal second, so the units agree in the equation.

**Summary**

- A quantum relates to the energy gained or absorbed by an object.
- The value of Planck's constant is \(h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}\).

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