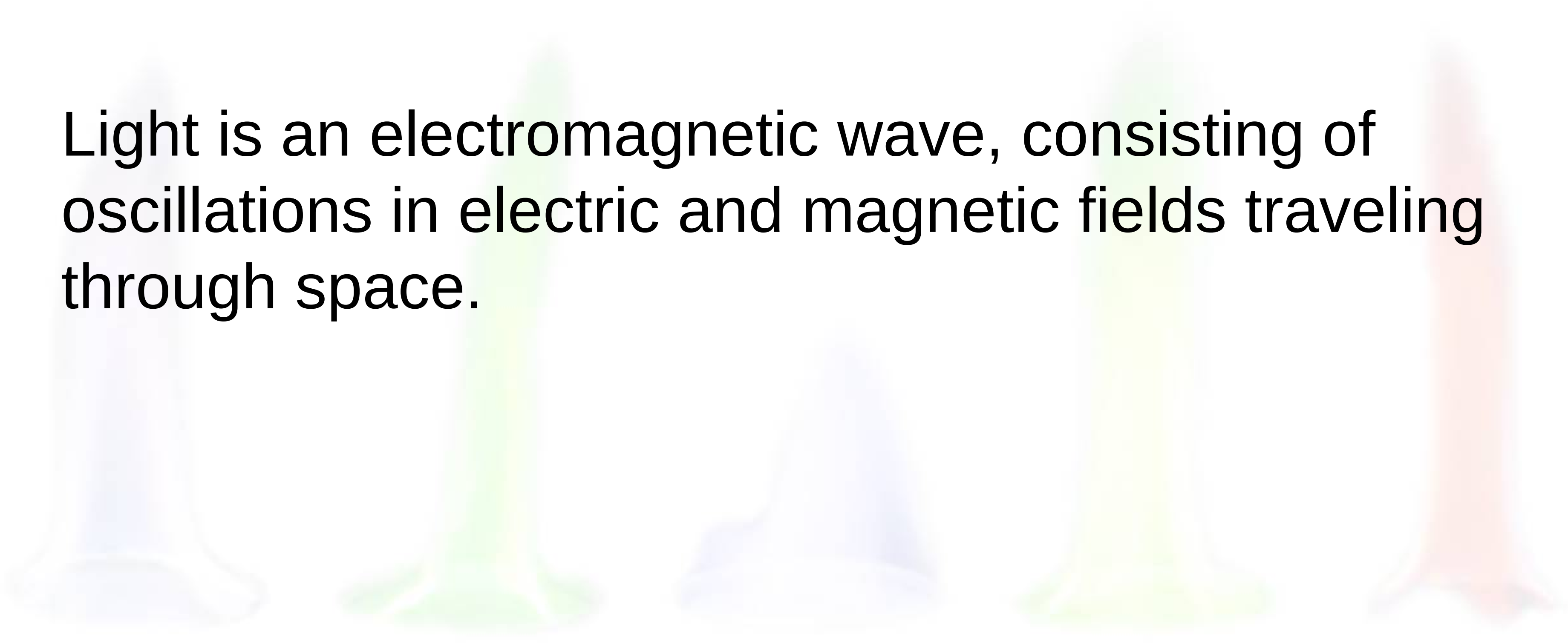
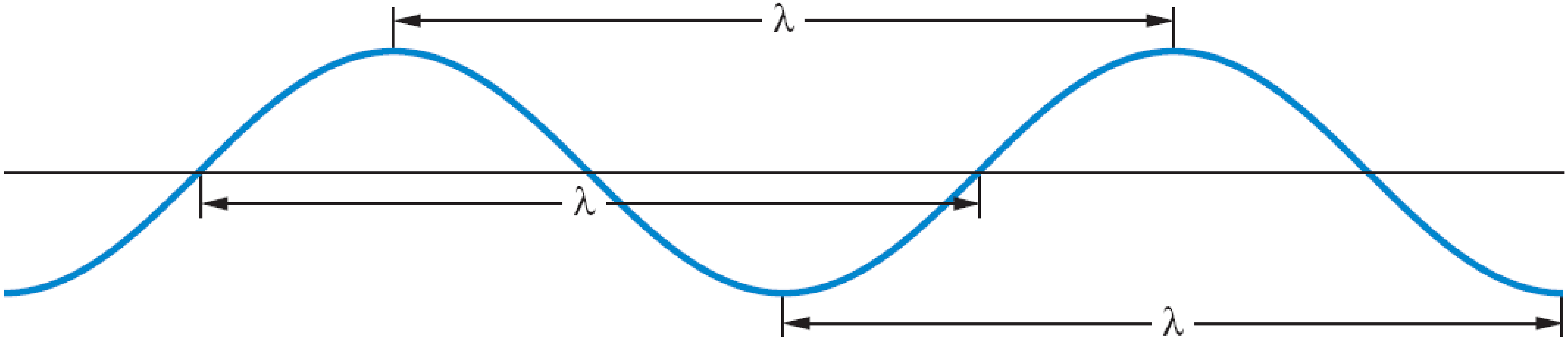


A **wave** is a continuously repeating change or oscillation in matter or in a physical field.

Light is an electromagnetic wave, consisting of oscillations in electric and magnetic fields traveling through space.

A wave can be characterized by its wavelength and frequency.

Wavelength, symbolized by the Greek letter lambda, λ, is the distance between any two identical points on adjacent waves.

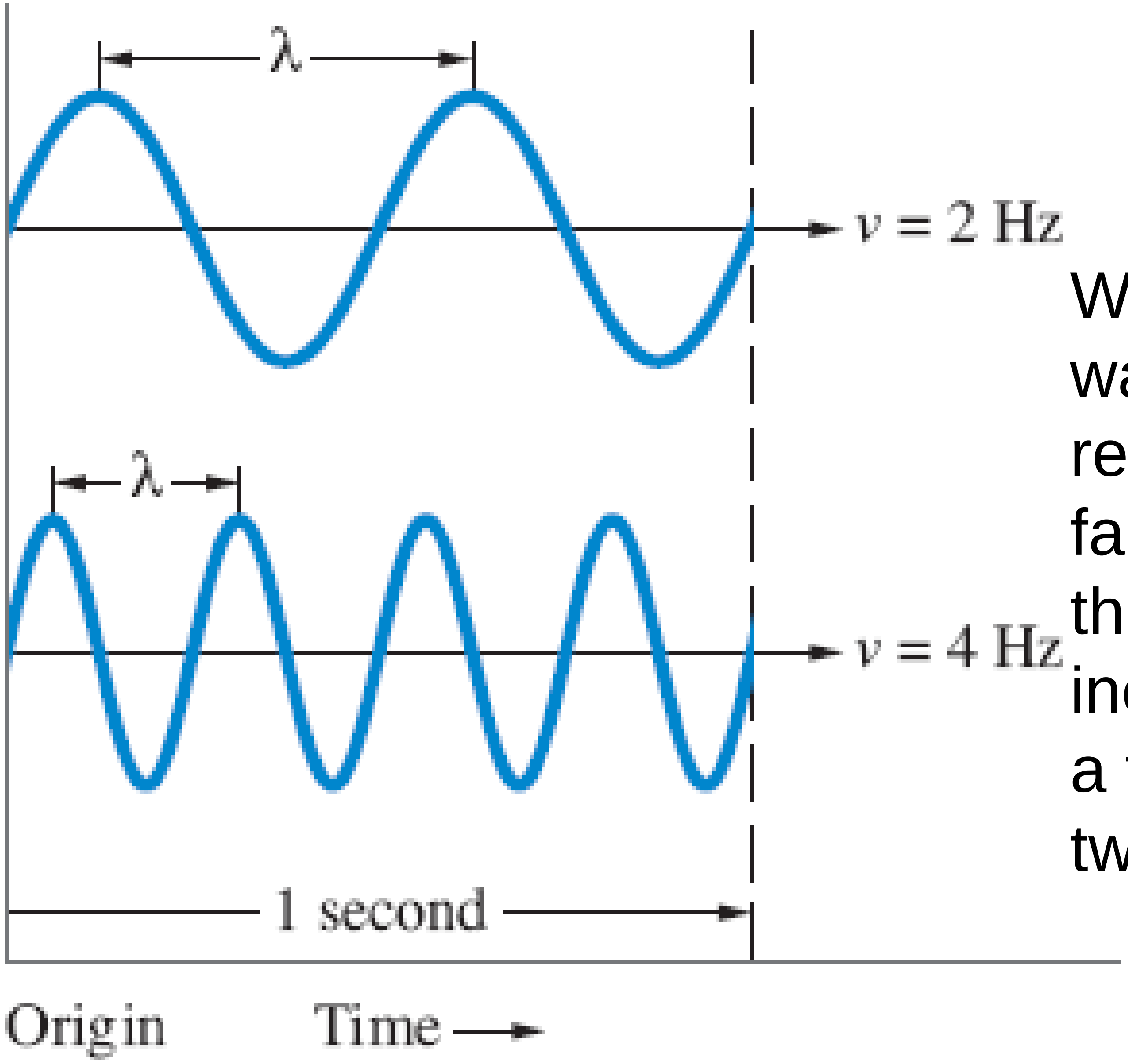


Frequency, symbolized by the Greek letter nu, ν, is the number of wavelengths that pass a fixed point in one unit of time (usually a second). The unit is $^{1}/_{s}$ or s^{-1}, which is also called the Hertz (Hz).

Wavelength and frequency are related by the wave speed, which for light is *c*, the speed of light, 3.00 x 10^8 m/s.

$$c = \nu\lambda$$

The relationship between wavelength and frequency due to the constant velocity of light is illustrated on the next slide.



When the wavelength is reduced by a factor of two, the frequency increases by a factor of two.

What is the wavelength of blue light with a frequency of 6.4×10^{14}/s?

$\nu = 6.4 \times 10^{14}/\text{s}$

$c = 3.00 \times 10^{8}$ m/s

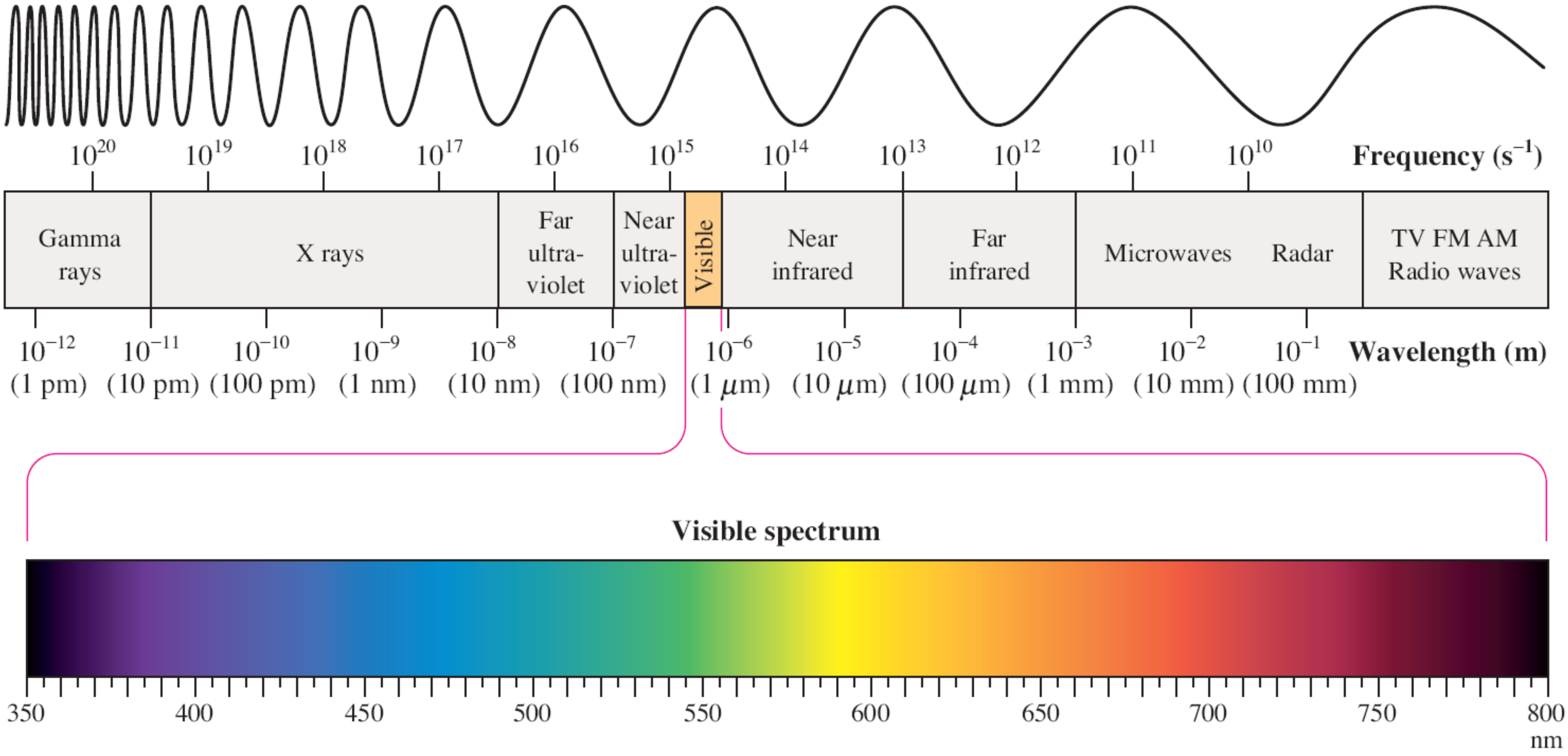
$c = \nu\lambda$ so

$\lambda = c/\nu$

$$\lambda = \frac{c}{\nu} = \frac{3.00 \text{ x } 10^{8}\ \dfrac{\text{m}}{\cancel{\text{s}}}}{6.4 \text{ x } 10^{14}\ \dfrac{1}{\cancel{\text{s}}}}$$

$$\boxed{\lambda = 4.7 \times 10^{-7}\text{ m}}$$

The range of frequencies and wavelengths of electromagnetic radiation is called the electromagnetic spectrum.



One property of waves is that they can be diffracted—that is, they spread out when they encounter an obstacle about the size of the wavelength.

In 1801, Thomas Young, a British physicist, showed that light could be diffracted. By the early 1900s, the wave theory of light was well established.

Quantum Effects and Photons

Planck’s Quantization of Energy (1900)

- According to Max Planck, the atoms of a solid oscillate with a definite frequency, ν.
- He proposed that an atom could have only certain energies of vibration, E, those allowed by the formula

$$E = h\nu$$

where h (Planck’s constant) is assigned a value of 6.63×10^{-34} J· s and n must be an integer.

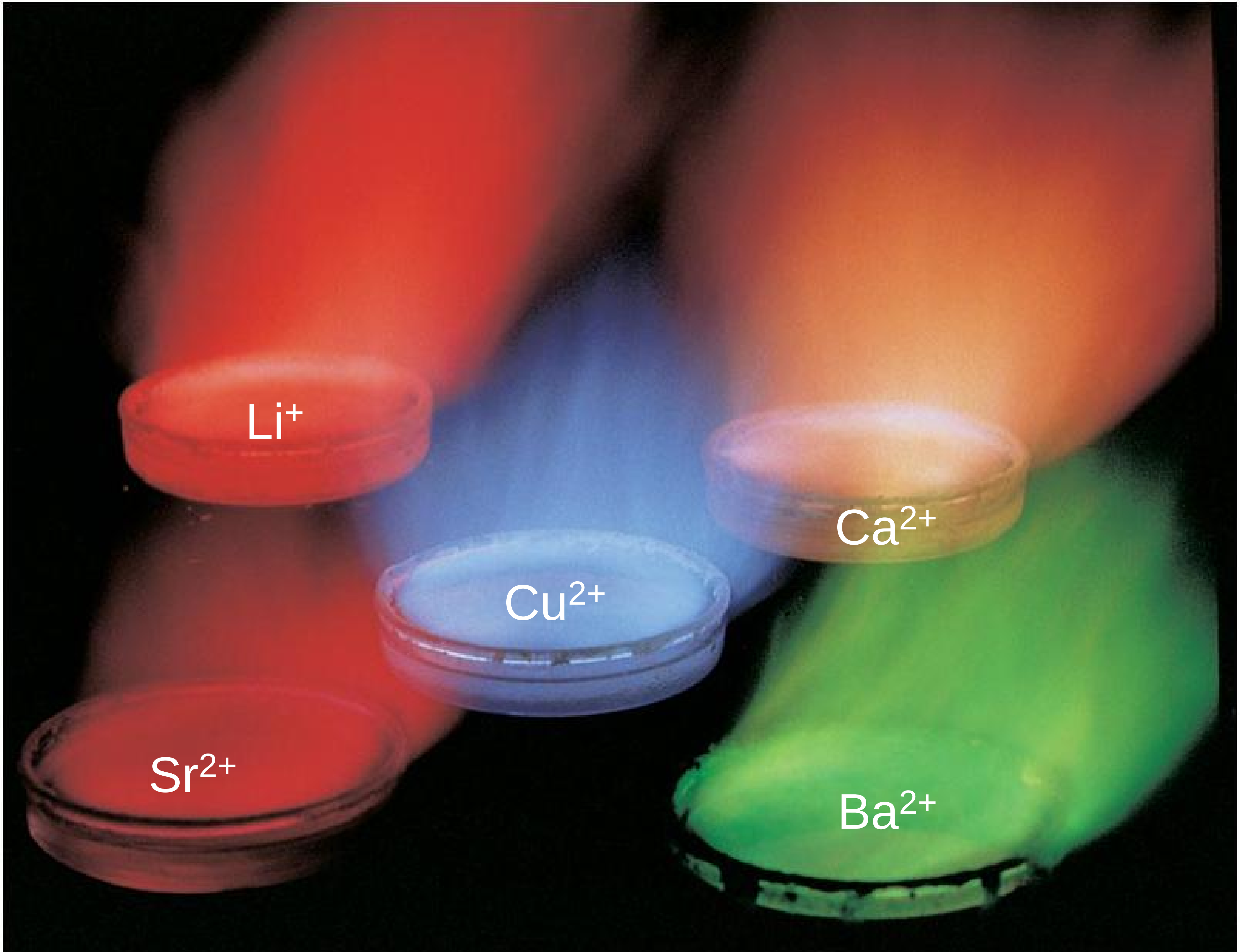
Failures of Classial Physics

Line spectra of atoms

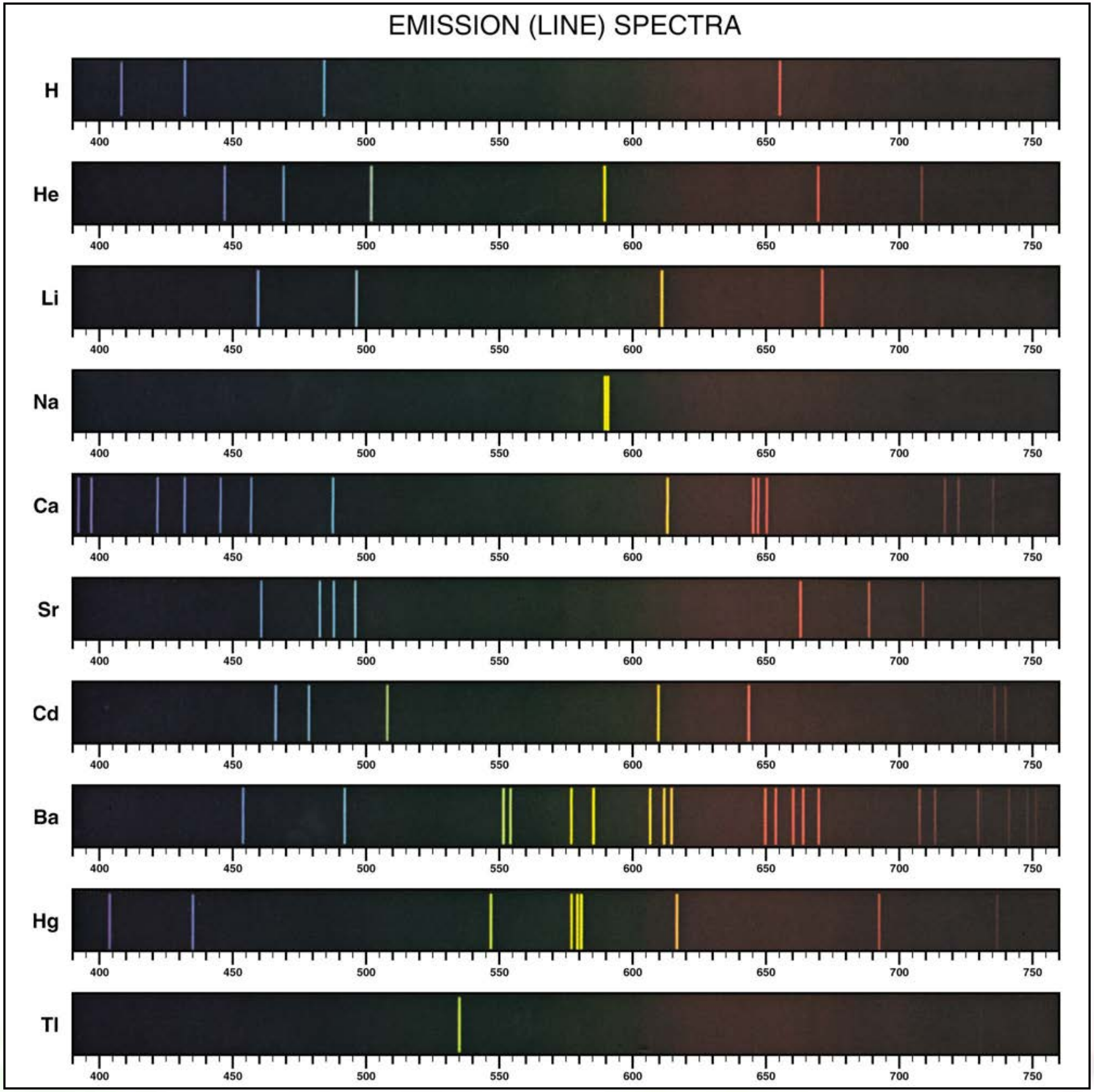
Black body radiation

Heat capacity of solids

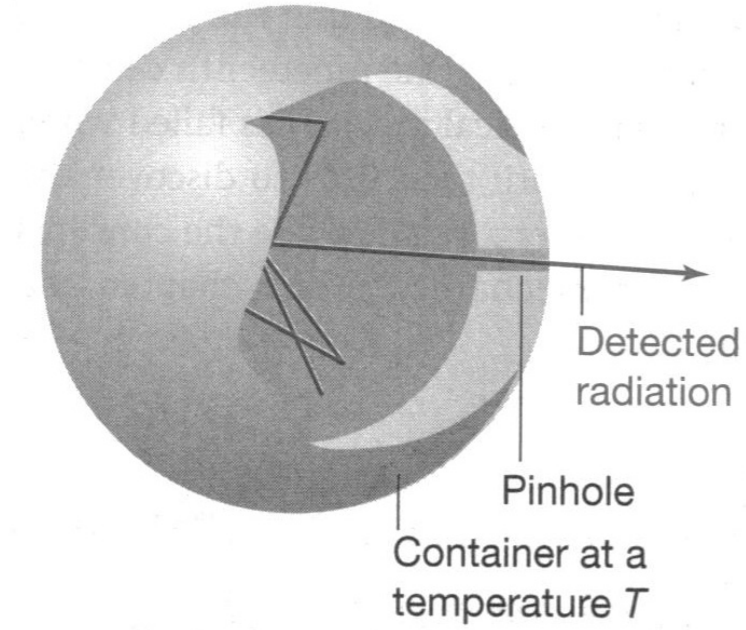
Photoelectric effect

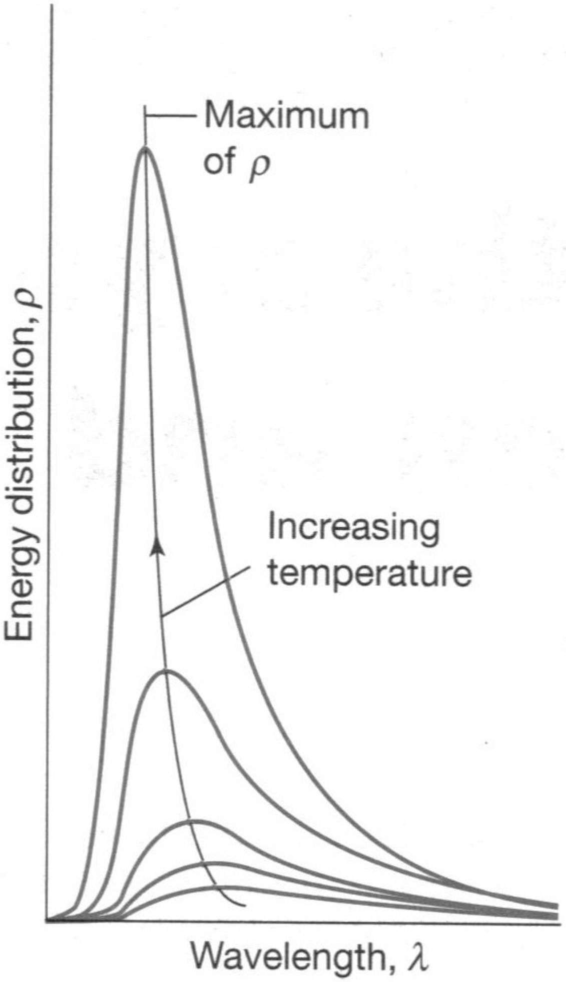

Li^+
Ca^{2+}
Cu^{2+}
Sr^{2+}
Ba^{2+}

Emission (line) spectra of some elements.

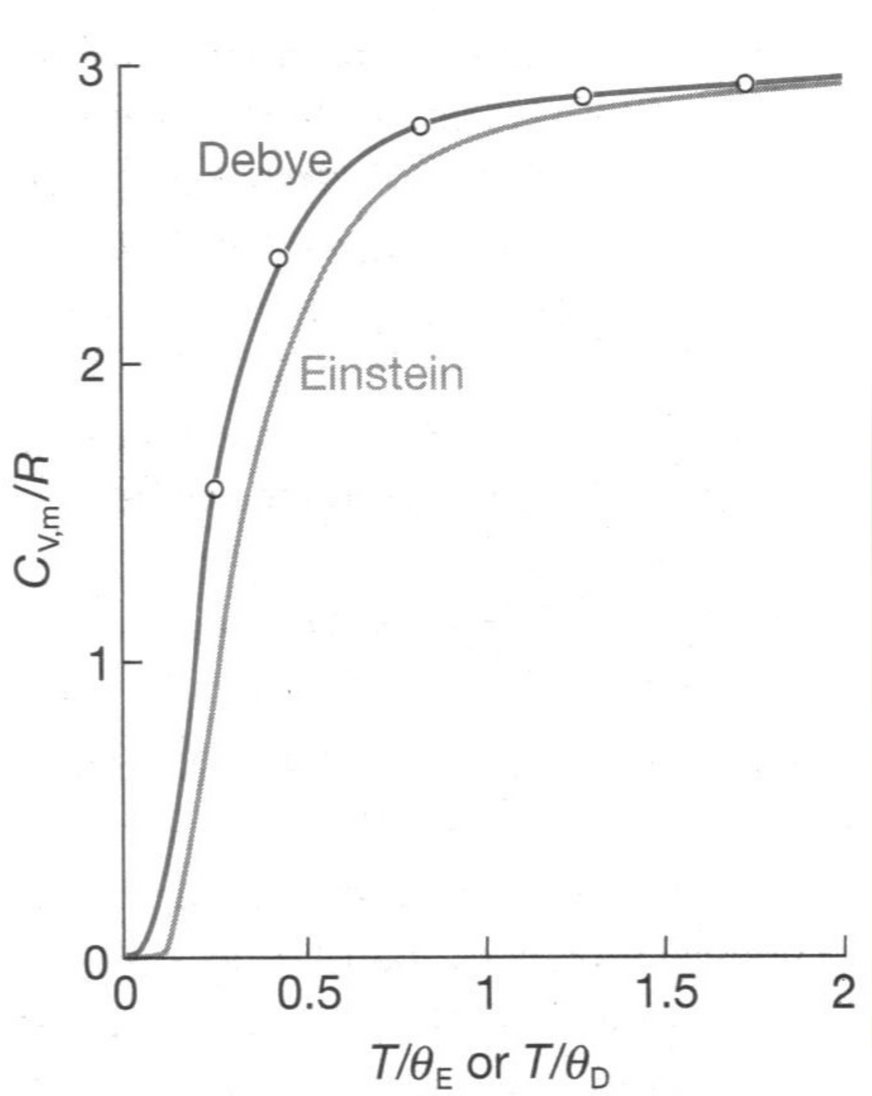


Blackbody radiation





Heat capacity of metals



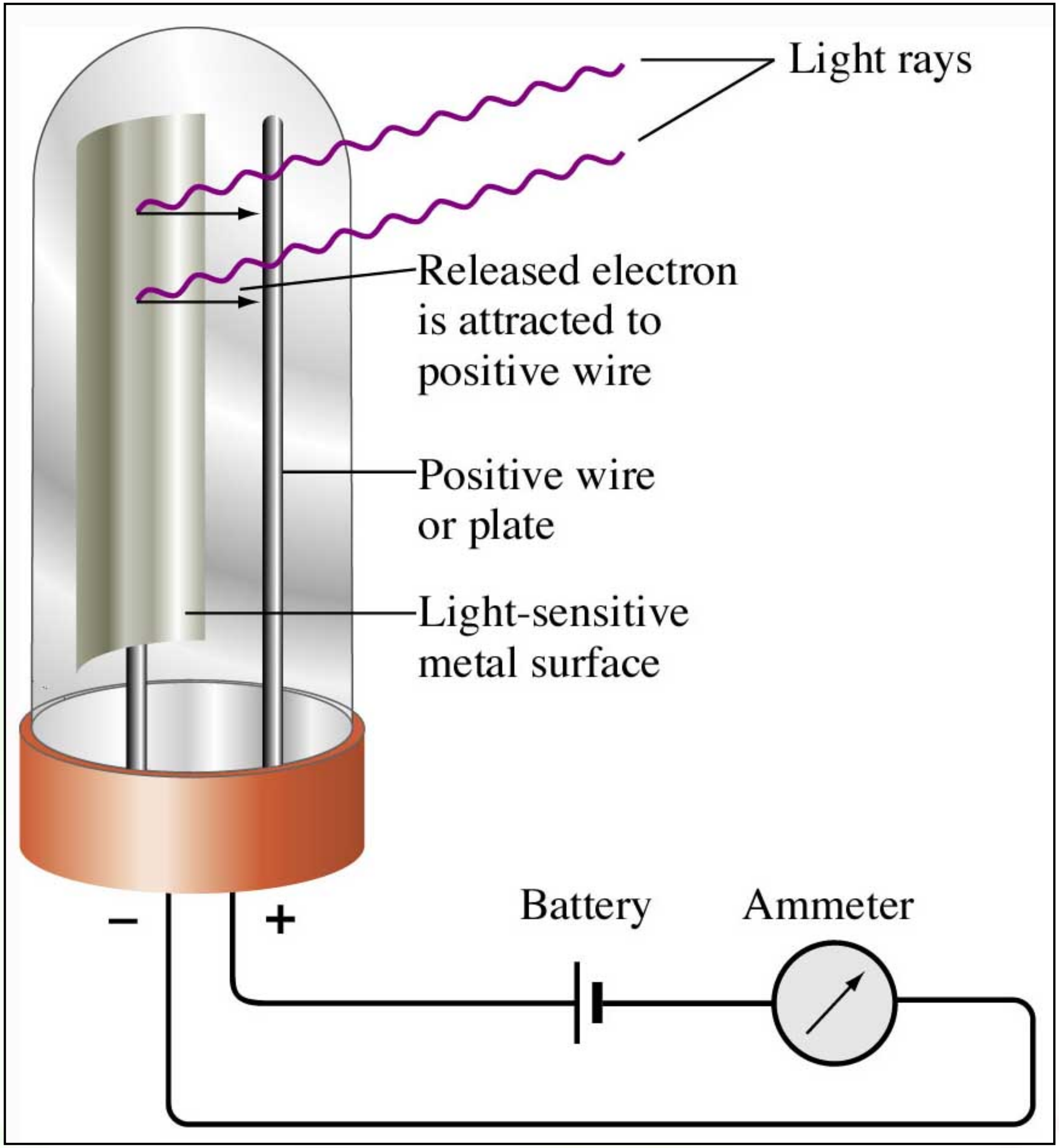
Quantum Effects and Photons

Planck’s Quantization of Energy.

- The only energies a vibrating atom can have are hν, 2hν, 3hν, and so forth.

- The numbers symbolized by n are quantum numbers.
- The vibrational energies of the atoms are said to be quantized.
- Solved the ultraviolet catastrophe in blackbody radiation

The photoelectric effect.



Quantum Effects and Photons

Photoelectric Effect

- The photoelectric effect is the ejection of electrons from the surface of a metal when light shines on it.
- Electrons are ejected only if the light exceeds a certain “threshold” frequency.
- Violet light, for example, will cause potassium to eject electrons, but <u>no amount of red light</u> (which has a lower frequency) has <u>any</u> effect.

Quantum Effects and Photons

By the early part of twentieth century, the wave theory of light seemed to be well entrenched.

- In 1905, Albert Einstein proposed that light had both wave and particle properties to explain the observations in the photoelectric effect.
- Einstein based this idea on the work of a German physicist, Max Planck.

Quantum Effects and Photons

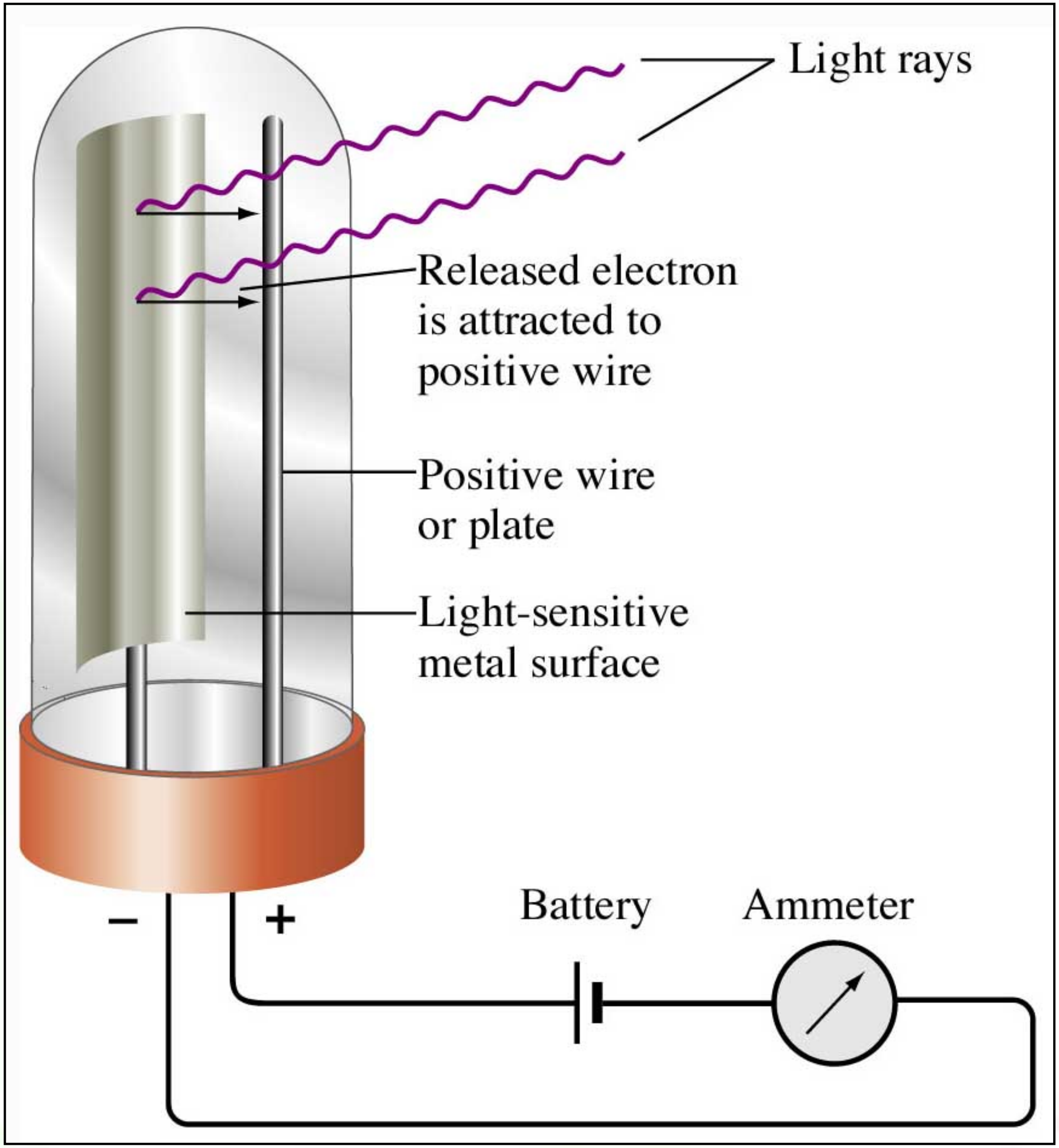
Photoelectric Effect

- The energy of the photons proposed by Einstein would be proportional to the observed frequency, and the proportionality constant would be Planck’s constant.

$$E = h\nu$$

- In 1905, Einstein used this concept to explain the “photoelectric effect.”

The photoelectric effect.



Quantum Effects and Photons

Photoelectric Effect

- Einstein’s assumption that an electron is ejected when struck by a single photon implies that it behaves like a particle.
- When the photon hits the metal, its energy, $h\nu$ is taken up by the electron.
- The photon ceases to exist as a particle; it is said to be “absorbed.”

Quantum Effects and Photons

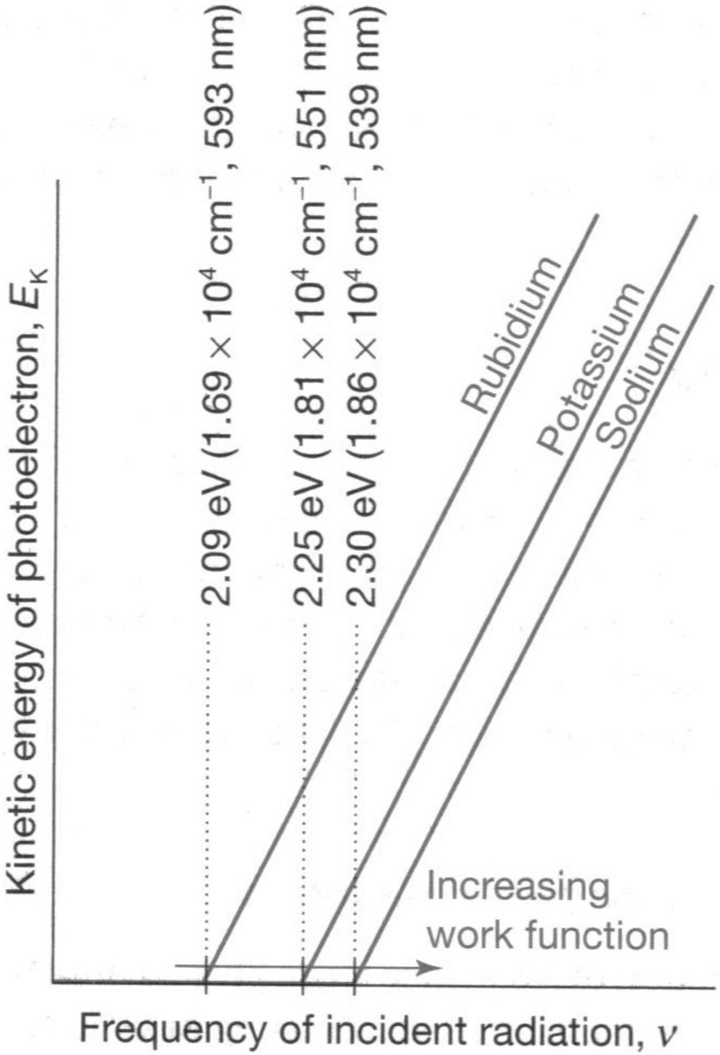
Photoelectric Effect

- The “wave” and “particle” pictures of light should be regarded as complementary views of the same physical entity.

- This is called the wave-particle duality of light.
- The equation E = hν displays this duality; E is the energy of the “particle” photon, and ν is the frequency of the associated “wave.”

Photoelectric effect

- For a given metal, a certain amount of energy is needed to eject the electron
- This is called the work function
- Since E=hν, the photons must have a frequency higher than the work function in order to eject electrons

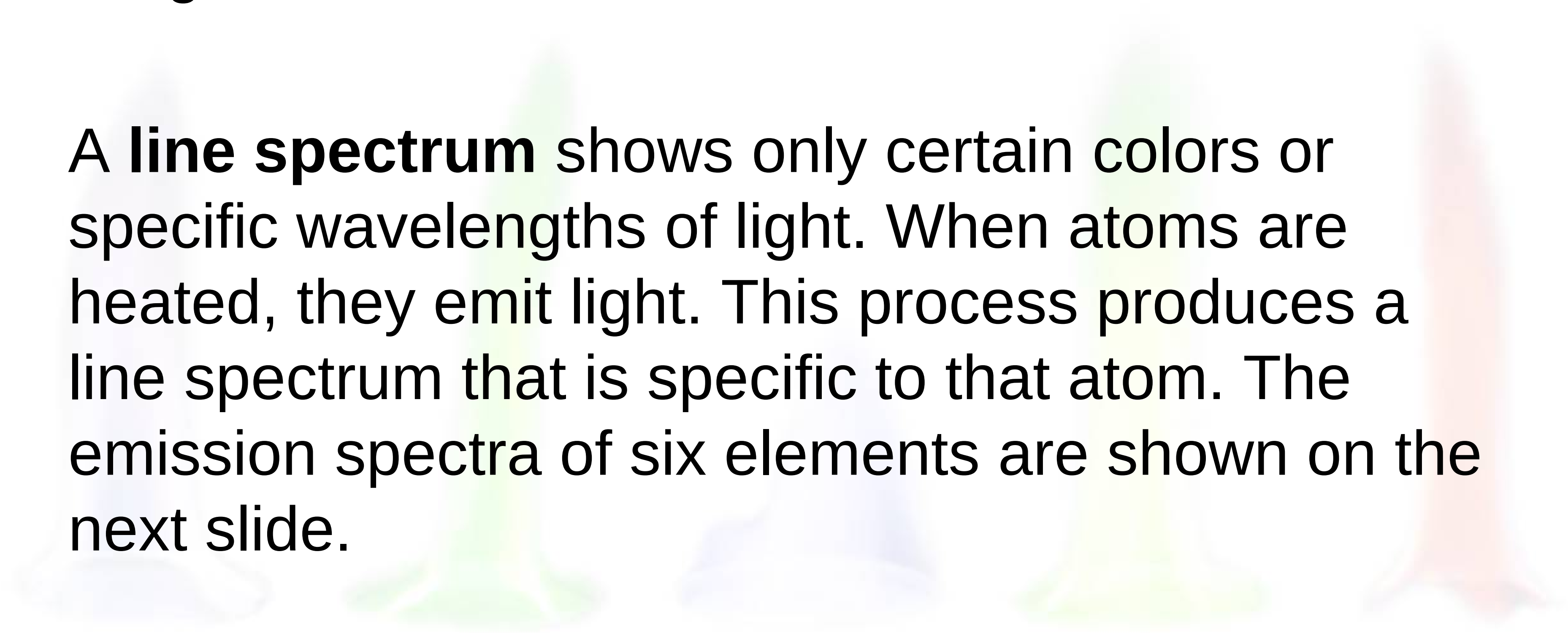


In the early 1900s, the atom was understood to consist of a positive nucleus around which electrons move (Rutherford’s model).

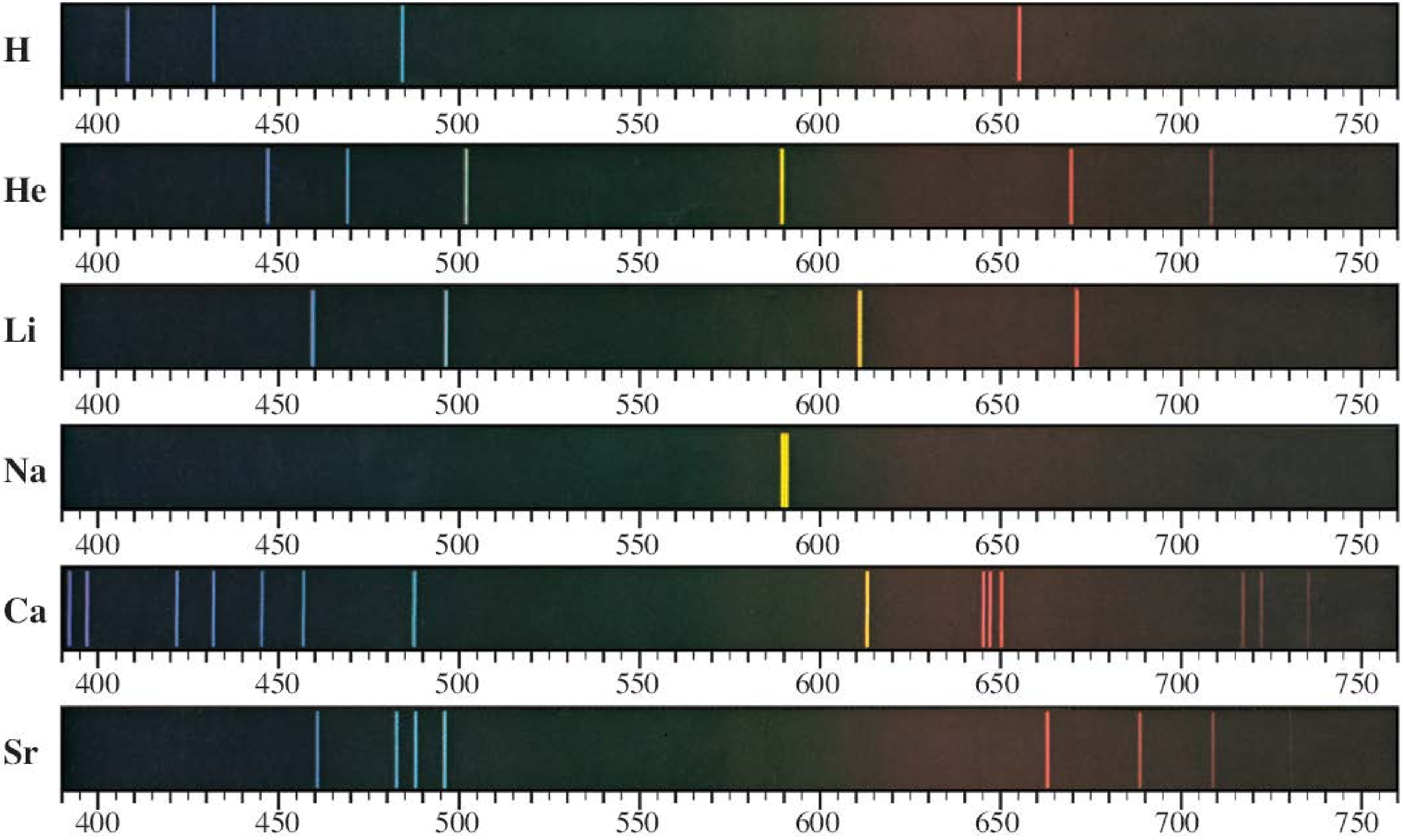
This explanation left a theoretical dilemma: According to the physics of the time, an electrically charged particle circling a center would continually lose energy as electromagnetic radiation. But this is not the case—atoms are stable.

In addition, this understanding could not explain the observation of line spectra of atoms.

A **continuous spectrum** contains all wavelengths of light.



A **line spectrum** shows only certain colors or specific wavelengths of light. When atoms are heated, they emit light. This process produces a line spectrum that is specific to that atom. The emission spectra of six elements are shown on the next slide.


H
He
Li
Na
Ca
Sr
400
450
500
550
600
650
700
750

In 1913, Neils Bohr, a Danish scientist, set down postulates to account for

1. The stability of the hydrogen atom
2. The line spectrum of the atom

Energy-Level Postulate

An electron can have only certain energy values, called energy levels. Energy levels are quantized.

For an electron in a hydrogen atom, the energy is given by the following equation:

$$E = -\frac{R_H}{n^2}$$

$R_H = 2.179 \times 10^{-18}$ J

n = principal quantum number

Transitions Between Energy Levels

An electron can change energy levels by absorbing energy to move to a higher energy level or by emitting energy to move to a lower energy level.

For a hydrogen electron the energy change is given by

$$\Delta E = E_\mathrm{f} - E_\mathrm{i}$$

$$\Delta E = -R_\mathrm{H}\left(\frac{1}{n_\mathrm{f}^{\,2}} - \frac{1}{n_\mathrm{i}^{\,2}}\right)$$

$R_\mathrm{H} = 2.179 \times 10^{-18}$ J, Rydberg constant

The energy of the emitted or absorbed photon is related to ΔE:

$$E_{\textbf{photon}} = \left|\Delta E_{\text{electron}}\right| = h\boldsymbol{\nu}$$

$$h = \text{Planck's constant}$$

We can now combine these two equations:

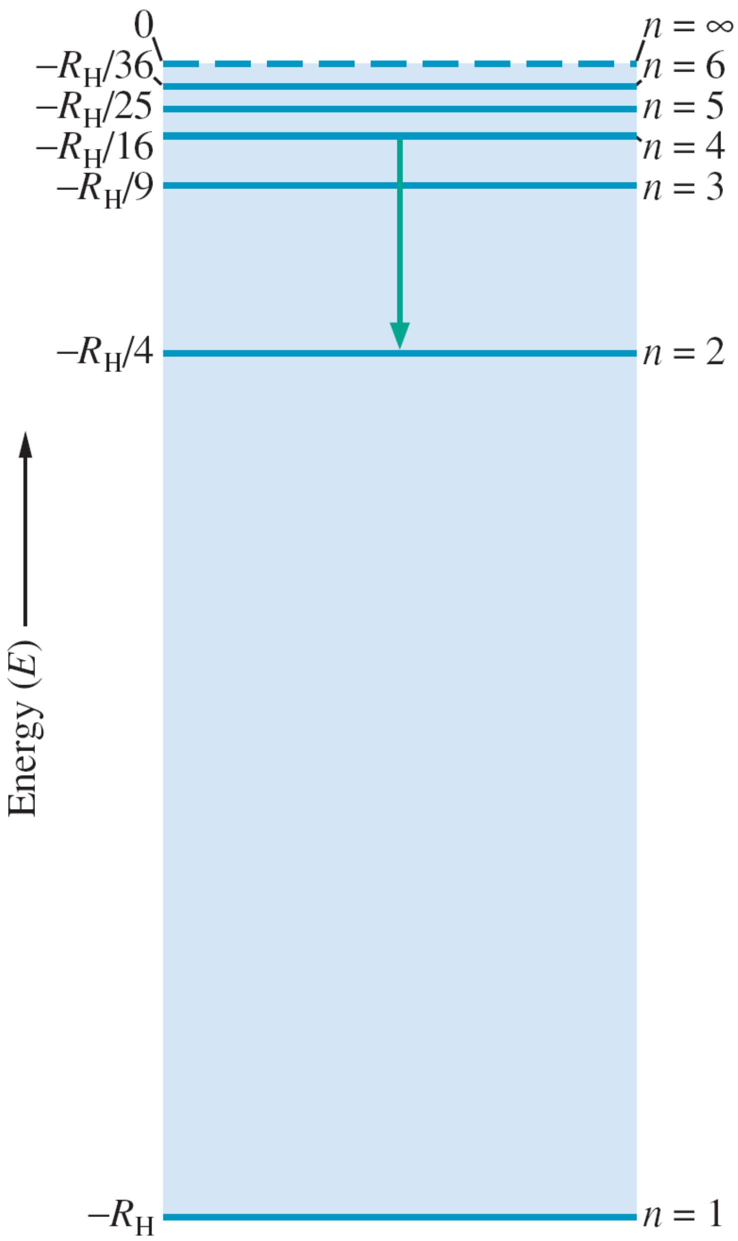
$$h\boldsymbol{\nu} = \left| -R_{\text{H}}\left(\frac{1}{{n_{\text{f}}}^{2}} - \frac{1}{{n_{\text{i}}}^{2}}\right)\right|$$

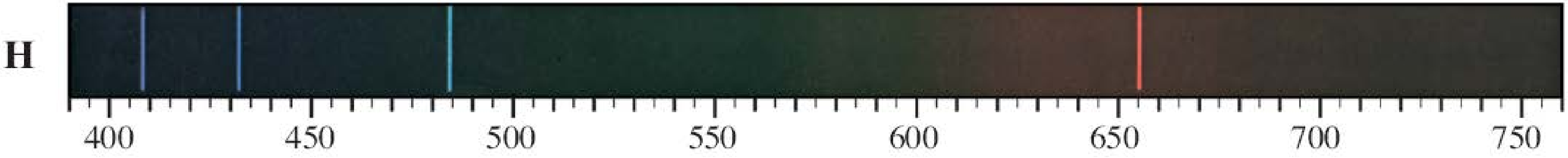
Light is **absorbed** by an atom when the electron transition is from lower n to higher n ($n_f > n_i$). In this case, ΔE will be positive.

Light is **emitted** from an atom when the electron transition is from higher n to lower n ($n_f < n_i$). In this case, ΔE will be negative.

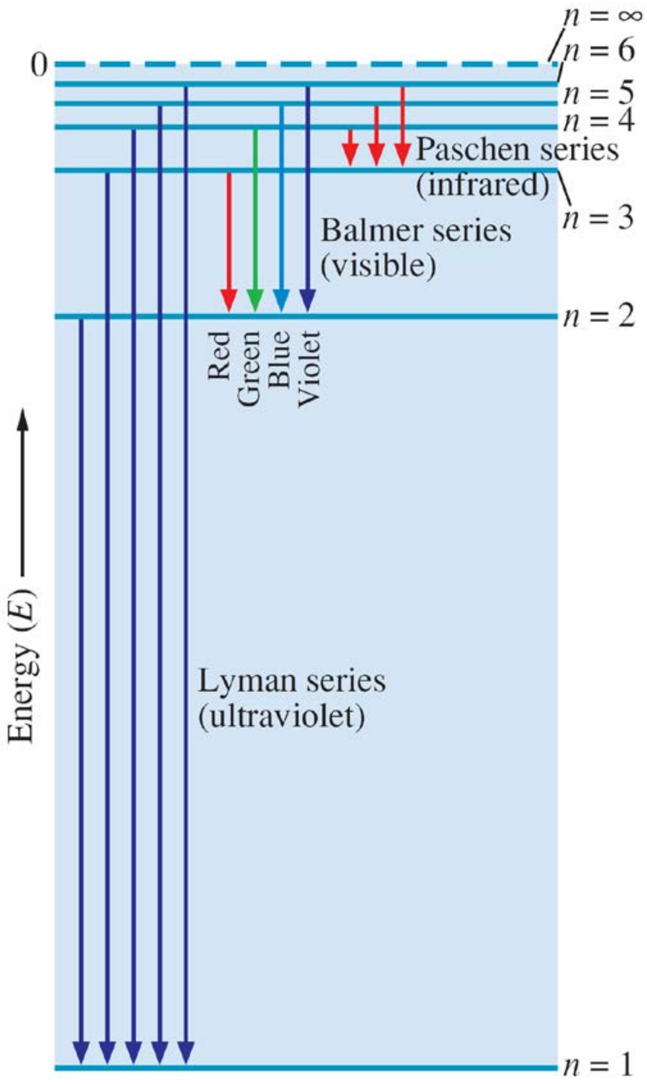
An electron is **ejected** when $n_f = \infty$.

Energy-level diagram for the hydrogen atom.





Electron transitions for an electron in the hydrogen atom.



Planck

Vibrating atoms have only certain energies:

$$E = h\nu \text{ or } 2h\nu \text{ or } 3h\nu$$

Einstein

Energy is quantized in particles called photons:

$$E = h\nu$$

Bohr

Electrons in atoms can have only certain values of energy. For hydrogen:

$$E = -\frac{R_H}{n^2}$$

$R_H = 2.179 \times 10^{-18}$ J, n = principal quantum number

X-ray spectra

When energetic electrons bombard a solid target, two kinds of x-rays are emitted:

1) Bremsstrahlung or “braking radiation” due to the fast electrons decelerating;

2) Characteristic line emission from the filling of inner electron vacancies (core holes) produced as a result of ionization of the atoms by the fast electrons.

The x-ray energy for an L electron making a transition to a half filled K shell in an atom with atomic number *Z* is approximately

$$\Delta E_K - E_L = (Z-1)^2 Ry\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = \tfrac{3}{4}(Z-1)^2 Ry$$

THE HIGH FREQUENCY SPECTRA OF THE ELEMENTS Moseley's law

By H. G. J. Moseley, M. A.

Phil. Mag. (1913), p. 1024

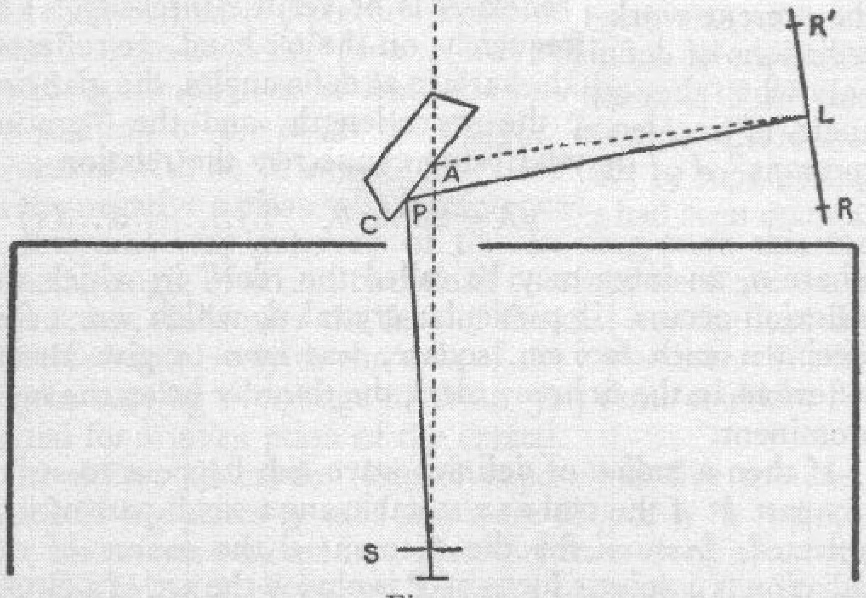


Fig. 1.

S – slit

C – ferric sulphide crystal

L – photographic plate

RR' – reference line on the plate

Notice that numbers 43 and 61 are Tc and Pm are radioactive and were unknown at the time!

High-Frequency Spectra of the Elements.

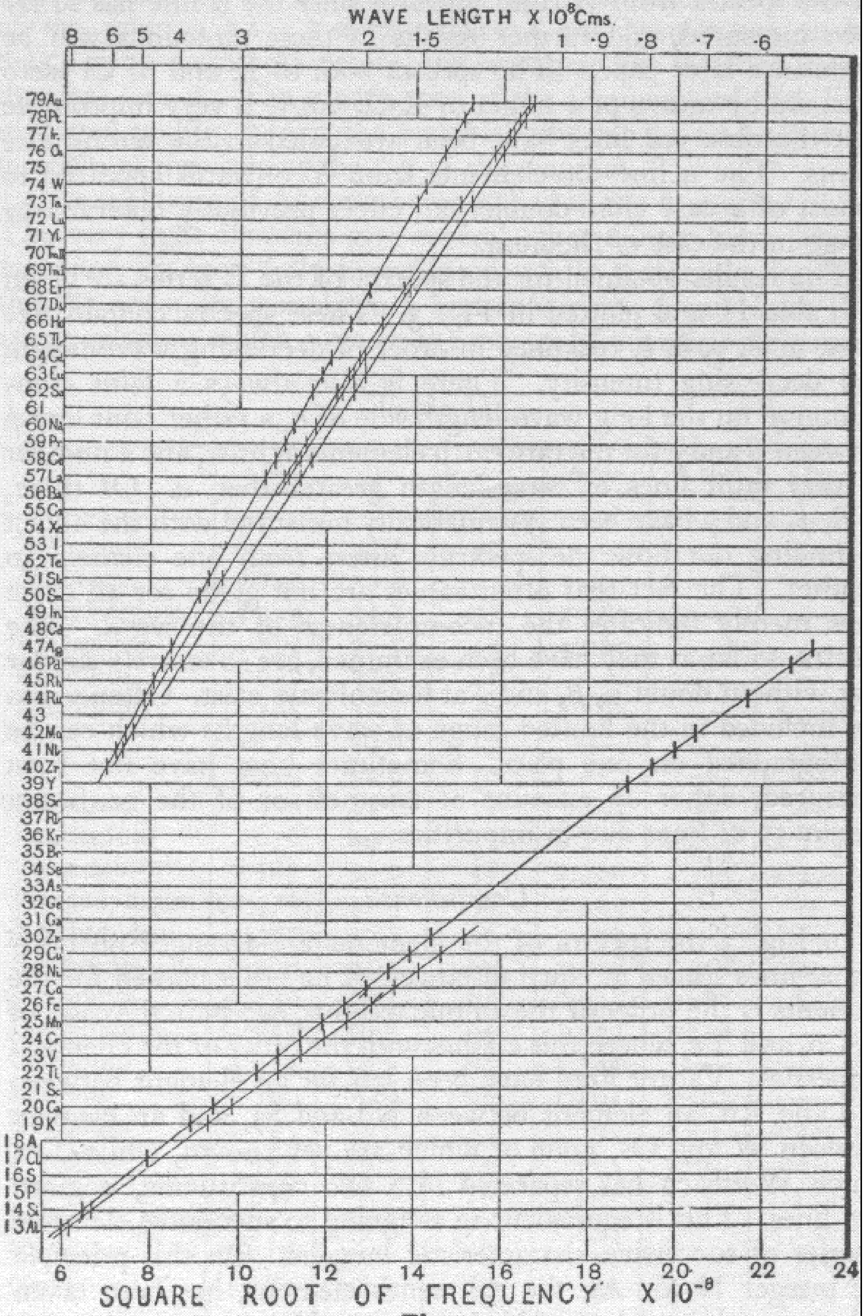


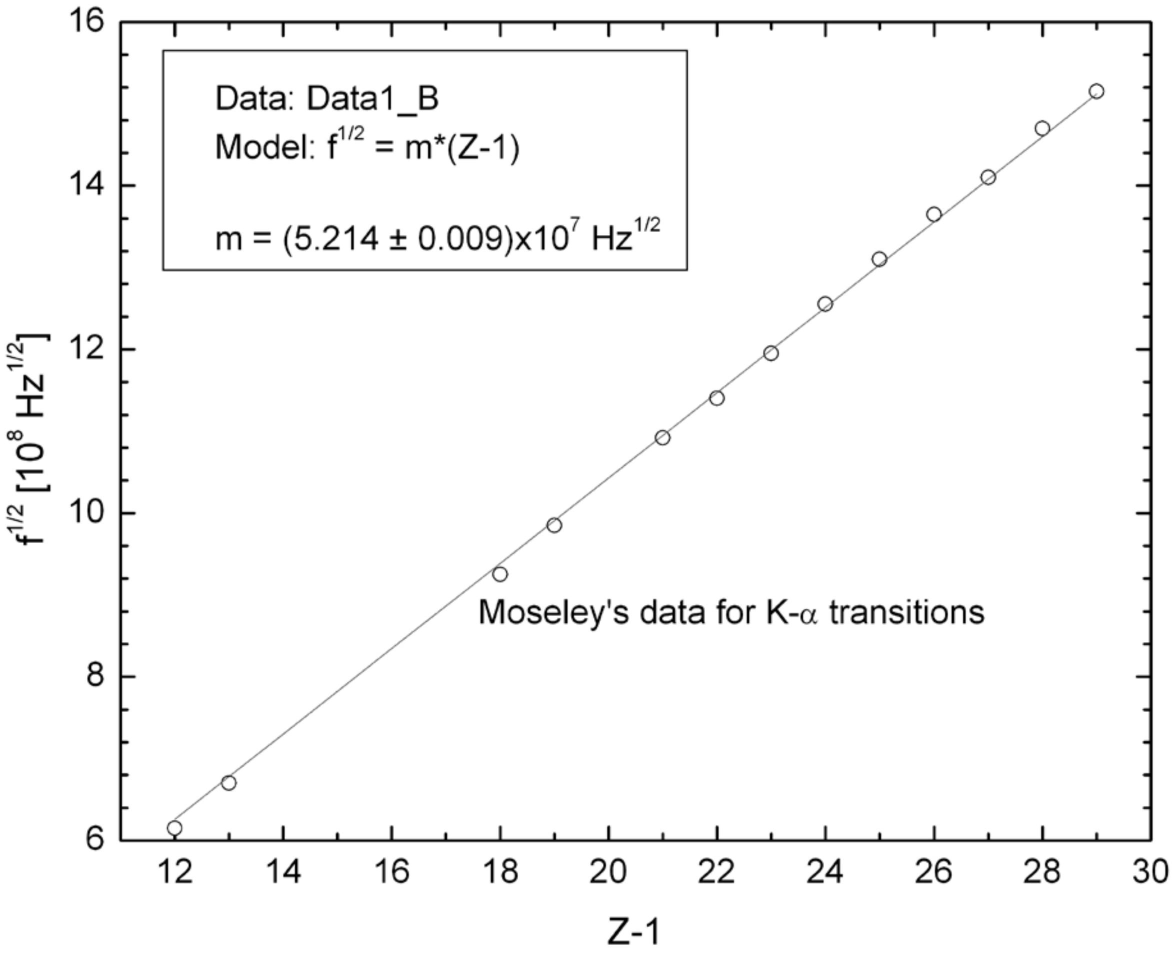
Fig. 3.

THE HIGH FREQUENCY SPECTRA OF THE ELEMENTS

By H. G. J. Moseley, M. A.
Phil. Mag. (1913), p. 1024

Moseley’s law

$$\sqrt{\tfrac{3}{4}Ry/h}\times(Z-1)=4.963\times10^{7}\sqrt{Hz}\times(Z-1)$$



Significance of Moseley’s work

Moseley found a systematic shift towards shorter wavelengths as one passed from one element to others of higher atomic weight, but there were some irregularities. To get over the difficulty posed by the irregularities, he assigned a number to each element, specifying its position in the periodic table. Then he could assign a relation between the frequency of X-ray lines and the atomic number - a relation known as Moseley's law.

When the elements were arranged according to the atomic numbers assigned by Moseley, some inconsistencies apparent in the Mendeleev table were removed. Thus Moseley was the first to arrange the elements in order of atomic number, rather than atomic weight, so he can be considered to be responsible for the present-day arrangement of the elements.

from A and B Scott Science History

Moseley’s measurements also proved that the nucleus held an integral number of elemental charges, thus placing the nuclear model of the atom on a firm foundation.

Moseley, the man

THE HIGH FREQUENCY SPECTRA OF THE ELEMENTS
By H. G. J. Moseley, M. A.
Phil. Mag. (1913), p. 1024

Henry Moseley (1887-1915): A British chemist who studied under Rutherford and brilliantly developed the application of X-ray spectra to study atomic structure; his discoveries resulted in a more accurate positioning of elements in the Periodic Table by closer determination of atomic numbers. Tragically for the development of science, Moseley was killed in action at Gallipoli (the Dardanelles campaign) in 1915.

Moseley and WWI

http://en.wikipedia.org/wiki/Battle_of_Gallipoli