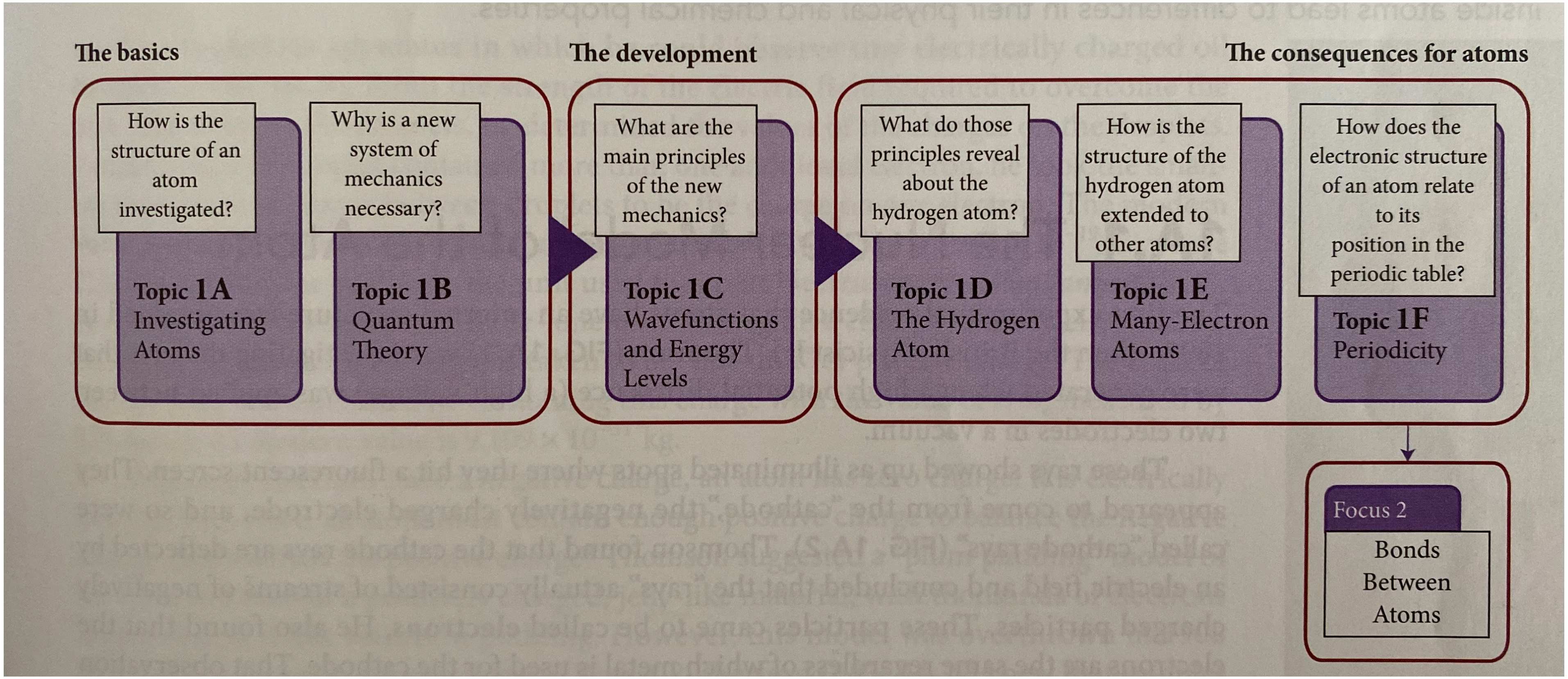
CH-110 Advanced General Chemistry I

Prof. A. Steinauer
angela.steinauer@epfl.ch

Investigating Atoms

Topic 1A

Overview Chapter 1 (Focus 1: Atoms)



Topic 1A.1: The nuclear model of the atom
Topic 1A.2: Electromagnetic radiation
Topic 1A.3: Atomic Spectra

WHY DO YOU NEED TO KNOW THIS MATERIAL?

- **Understanding of the structures of atoms is essential for understanding the differences in physical and chemical properties of substances**.
- Therefore: important to understand what is going on inside atoms and how their structure is studied.

WHAT DO YOU NEED TO KNOW ALREADY?

- Familiarity with the **nuclear model of the atom**: a small, positively charged nucleus surrounded by negatively charged electrons (Fundamentals B)

1A.1 The nuclear model of the atom

J.J. Thomson cathode ray experiment

First experimental evidence for internal structure of atom obtained in 1897 by British physicist **J. J. Thomson**.

- Cathode rays generated when a high potential difference (high voltage) was applied between two electrodes **in vacuum**:

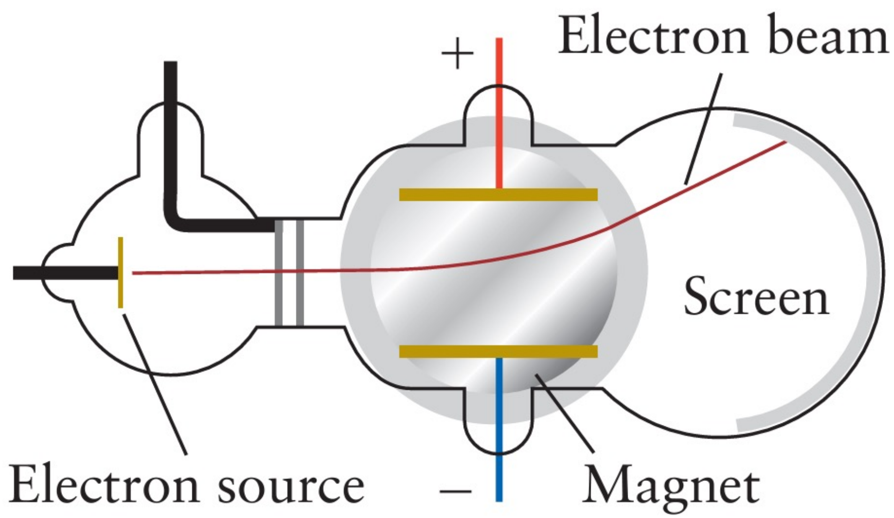


Figure 1A.2

1A.1 The nuclear model of the atom

J.J. Thomson cathode ray experiment

- Unknown particles reflected by electric field → must be negatively charged ("corpuscules") → later known as **electrons**
- Measured e/m_e: electron charge and mass of electron
- Same regardless which metal was used → must be part of all atoms: UNIVERSAL
- Challenged the long-held view of atoms as indivisible

1A.1 The nuclear model of the atom

Millikan oil drop apparatus

- Fine mist of oil droplets (atomizer)
- Parallel metal plates → uniform electric field
- Light source + microscope to observe droplets
- **X-rays** (not shown!) ionize air → some droplets gain electrons (become negatively charged)

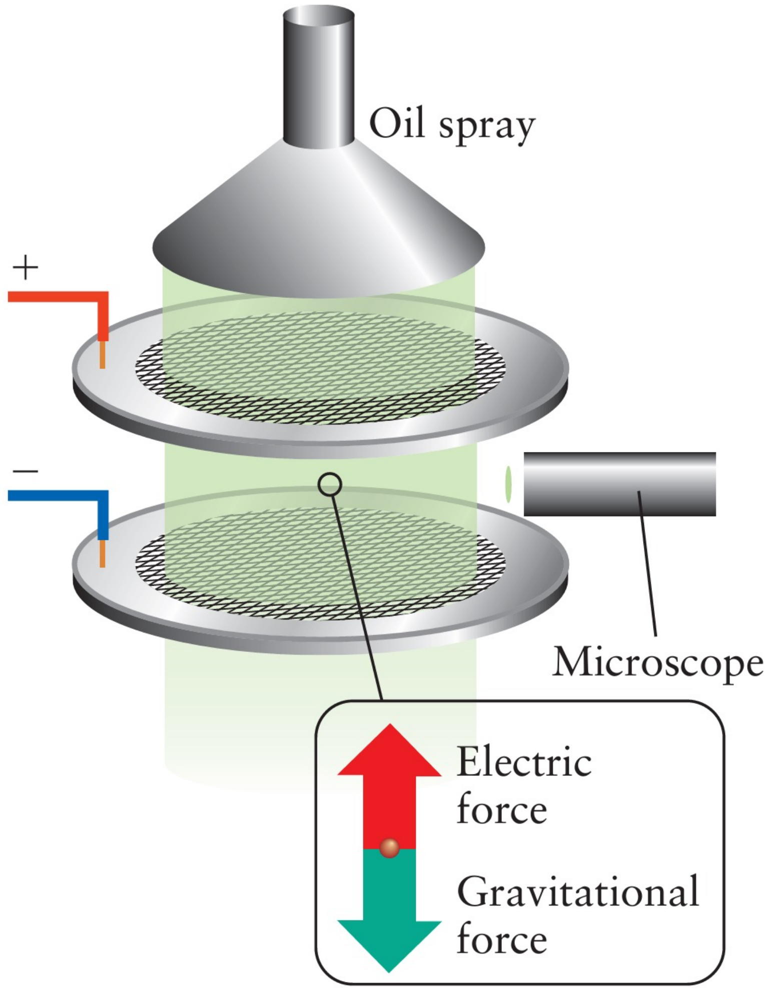


Figure 1A.3

1A.1 The nuclear model of the atom

Key findings

- Balanced electric force vs. gravity → measured droplet charge
- Repeated trials showed all charges were multiples of a single value
- Fundamental charge of the electron: **$e = 1.602 \times 10^{-19}$ C**

To practice: look at last year's exam!

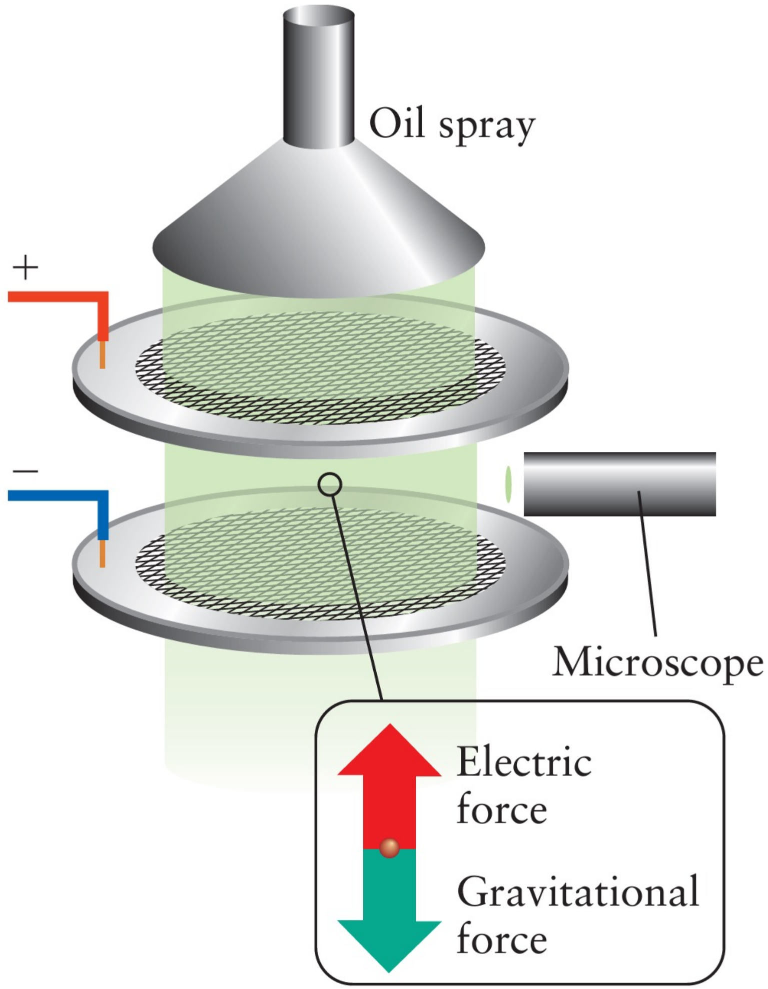


Figure 1A.3

1A.1 The nuclear model of the atom

What about the positive charge?

- Negative charge detectable, what about positive charge?
- **Plum pudding model** (Thomson):
- The atom was thought to be a relatively large, positively charged sphere with electrons scattered within it, balancing the charge.

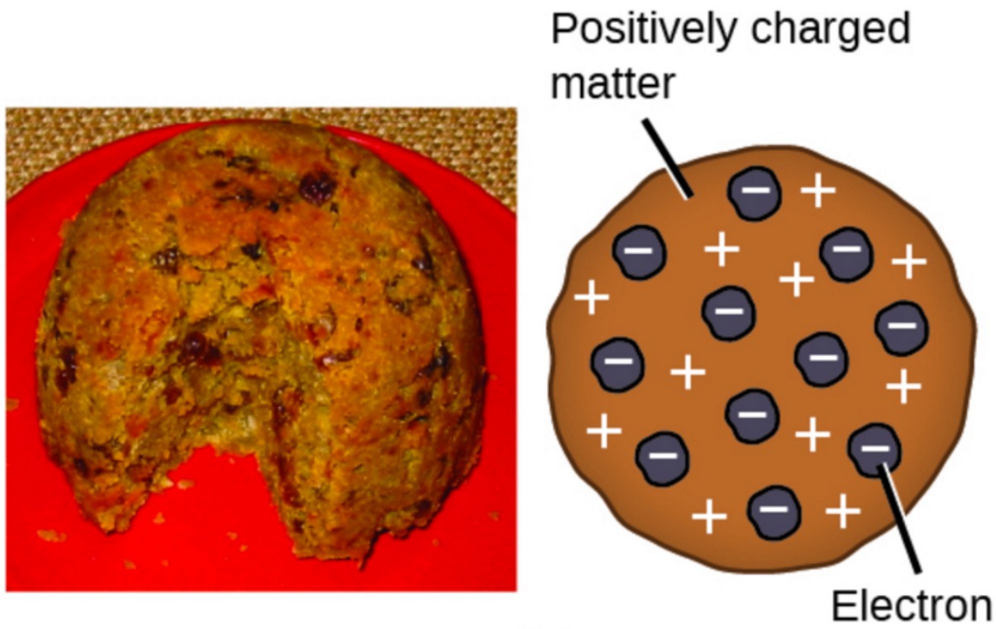


Image: https://www.khanacademy.org/science/chemistry/atomic-structure-and-properties/history-of-atomic-structure/a/discovery-of-the-electron-and-nucleus

1A.1 The nuclear model of the atom

The Rutherford gold foil experiment (Geiger-Marsden)

Rutherford discovered that some materials (e.g. Rn) emit positively charged particles, α (alpha) particles.

Alpha particles (α) shot at thin gold foil

Expectation: if positive charge spread out (plum pudding model) → only slight deflections

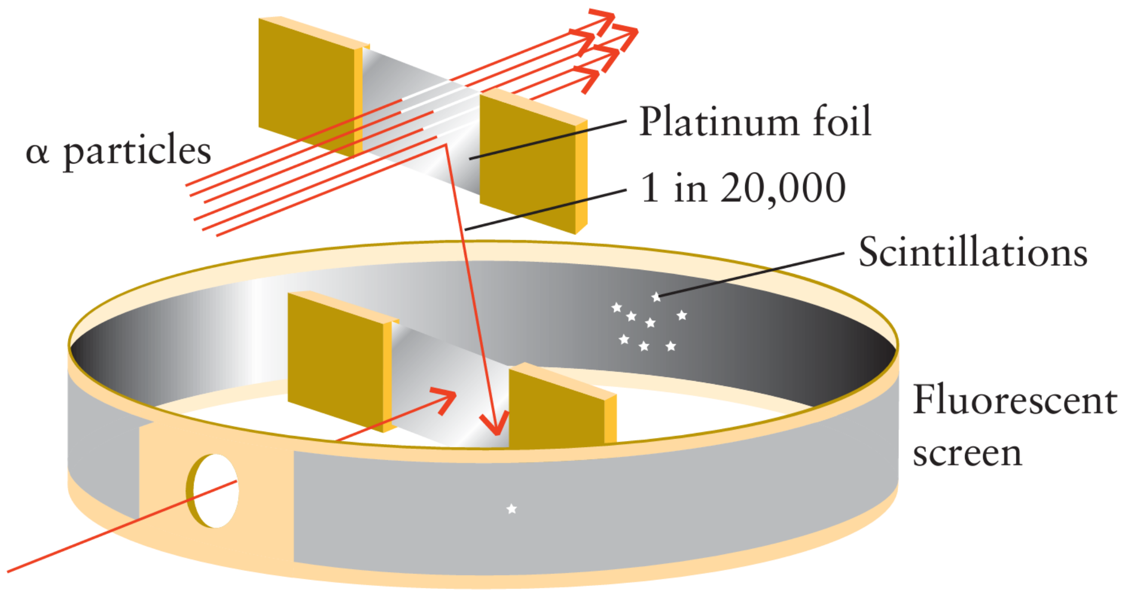


Figure 1A.5

Rutherford famously described the results as being as surprising as if you had **"fired a 15-inch shell at a piece of tissue paper and it came back and hit you."** The strong deflections indicated the presence of something very massive and compact within the atom.

Image: ChatGPT

1A.1 The nuclear model of the atom

The Rutherford gold foil experiment (Geiger-Marsden)

- Most α passed straight through → atom mostly empty space
- Few large deflections → concentrated positive charge in tiny, dense **nucleus**
- Electrons orbit nucleus at relatively large distances

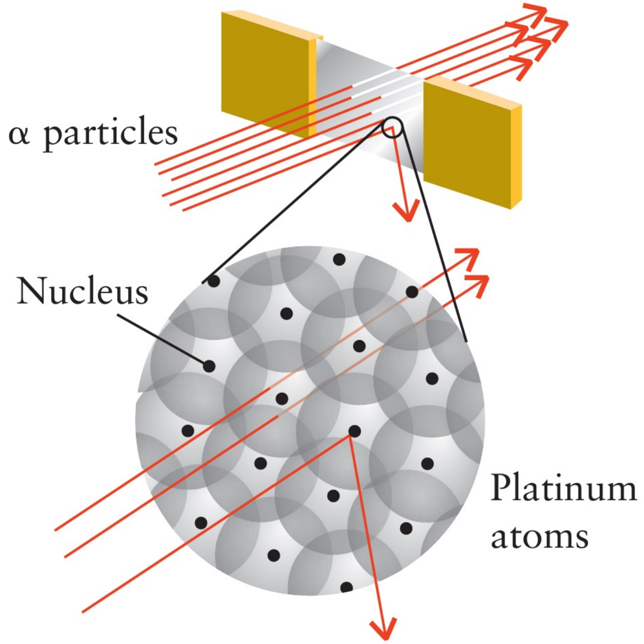


Figure 1A.6

1A.2 Electromagnetic radiation

How are the electrons arranged around the nucleus?

- Scientists used **light** to study this:
- They monitored the properties of light the atoms emit when stimulated by heat or an electric discharge.

→ **Spectroscopy:**

The study of how matter interacts with electromagnetic radiation (absorption, emission, scattering).

Applications: UV-Vis or fluorescence spectrophotometer, IR, NMR

1A.2 Electromagnetic radiation

Spectroscopy vs. spectrometry

Spectroscopy: Comes from **"spectrum"** (Latin for "appearance" or "image") and **"-scopy"** (Greek for "to observe"). It refers to the observation and study of the spectrum of light.

To remember:

Spectroscopy: **S** for "**S**tudying" the light spectrum.

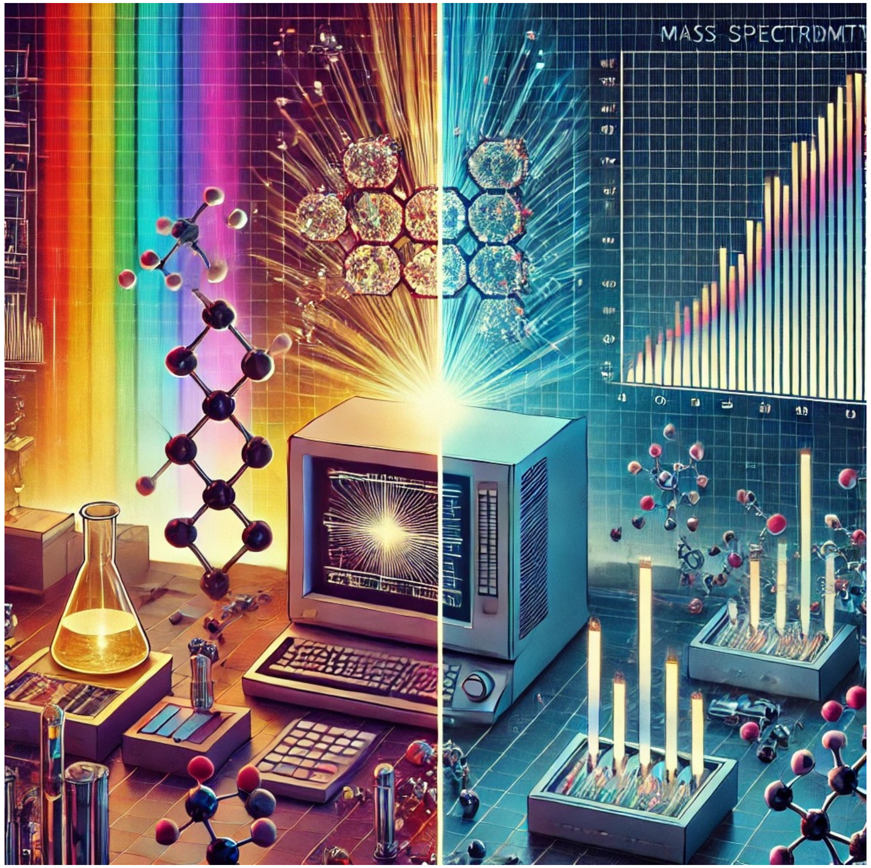


Image: ChatGPT

Spectrometry: Comes from **"spectrum"** and **"-metry"** (Greek for "to measure"). It refers to the measurement of properties of light or particles.

To remember:

Spectrometry: **M** for "**M**easuring" the light spectrum.

1A.2 Electromagnetic radiation

Light is a form of electromagnetic radiation

Image: ChatGPT

1A.2 Electromagnetic radiation

Light is a form of electromagnetic radiation

- Consists of **oscillating** (time-varying) **electric** and **magnetic** fields
- Travels through empty space at about 3 x 10^8 m s^{-1} (speed of light)
- Visible light, radio waves, microwaves, X-rays
- All forms of radiation **transfer energy** from one region of space to another.

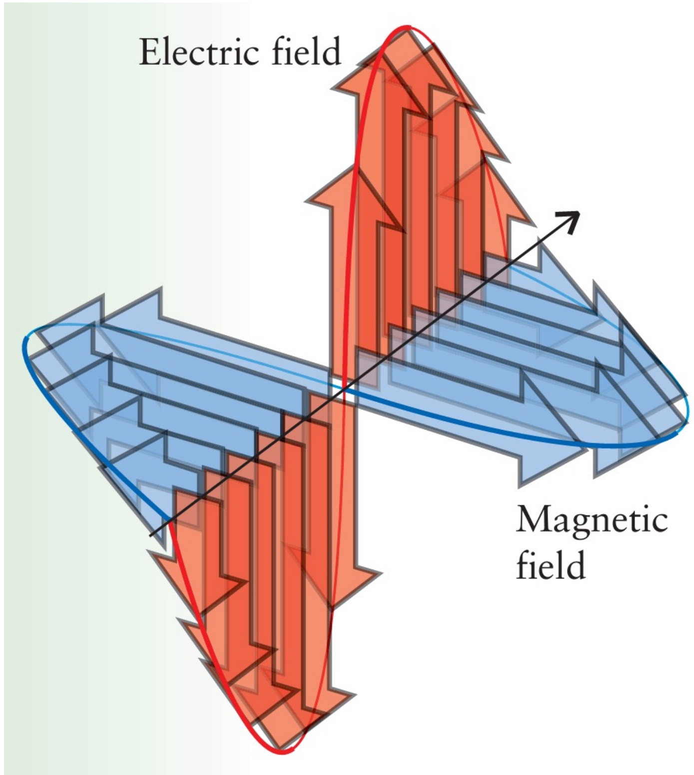


Figure A.8
(Fundamentals A)

1A.2 Electromagnetic radiation

Frequency of light

- Ray of light
- Electric field oscillates:

  (1) in direction

  (2) in strength

- Number of cycles per second:

  Frequency, *v* (Greek letter nu)

- Unit:

  Hertz (Hz), 1 Hz = 1 s^{-1}

Frequency of visible light radiation: 10^{15} Hz

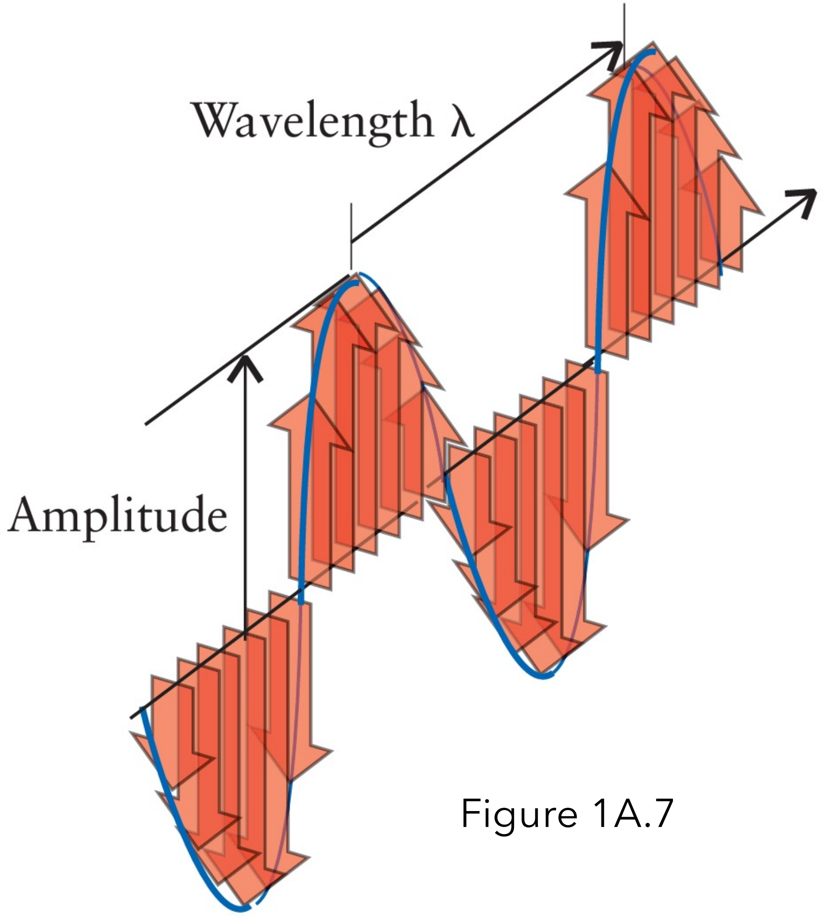


Figure 1A.7

1A.2 Electromagnetic radiation

Amplitude, intensity, and wavelength

- The **amplitude** is the height of the wave above the center line.
- The **square of the amplitude** is proportional to the **intensity**, or brightness, of the radiation.
- The **wavelength**, λ (the Greek letter lambda), is the peak-to-peak distance.

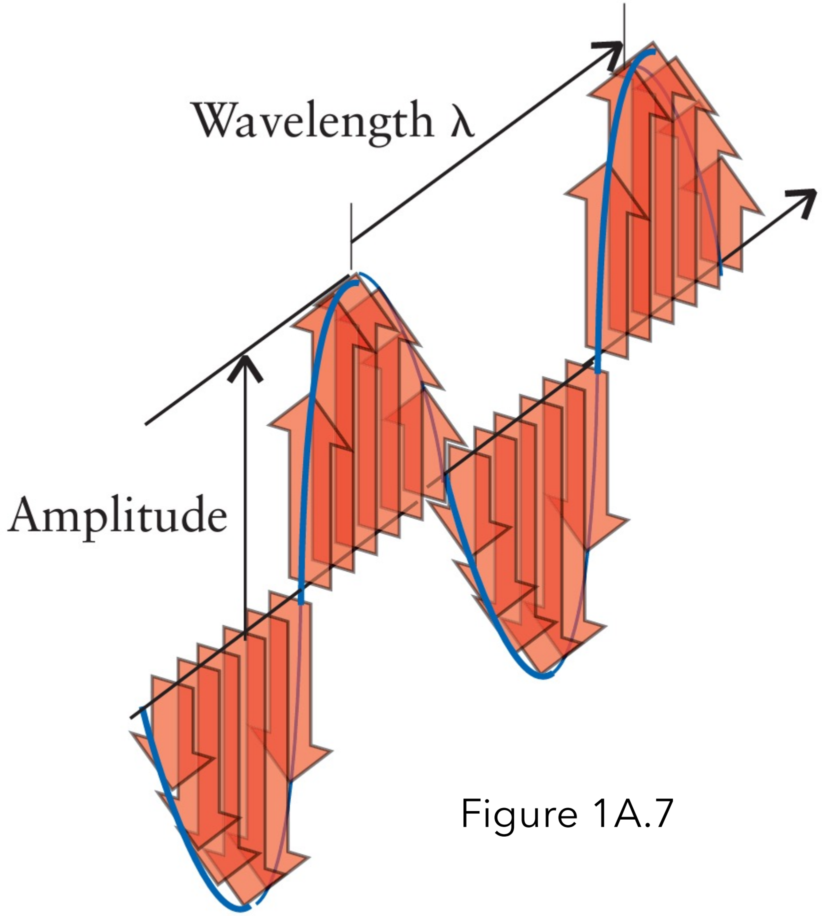


Figure 1A.7

$$\lambda \; x \; \nu = c$$

Important formula.
Need to know how to apply.

1A.2 Electromagnetic radiation

The wave equation for electromagnetic waves

$$\lambda \times \nu = c$$

- Expresses the relationship between the wavelength (λ), frequency (ν), and the speed of light (c).
- If the wavelength of light is short, many complete oscillations pass a given point in a second, if the wavelength is long, fewer complete oscillations pass the point in a second.

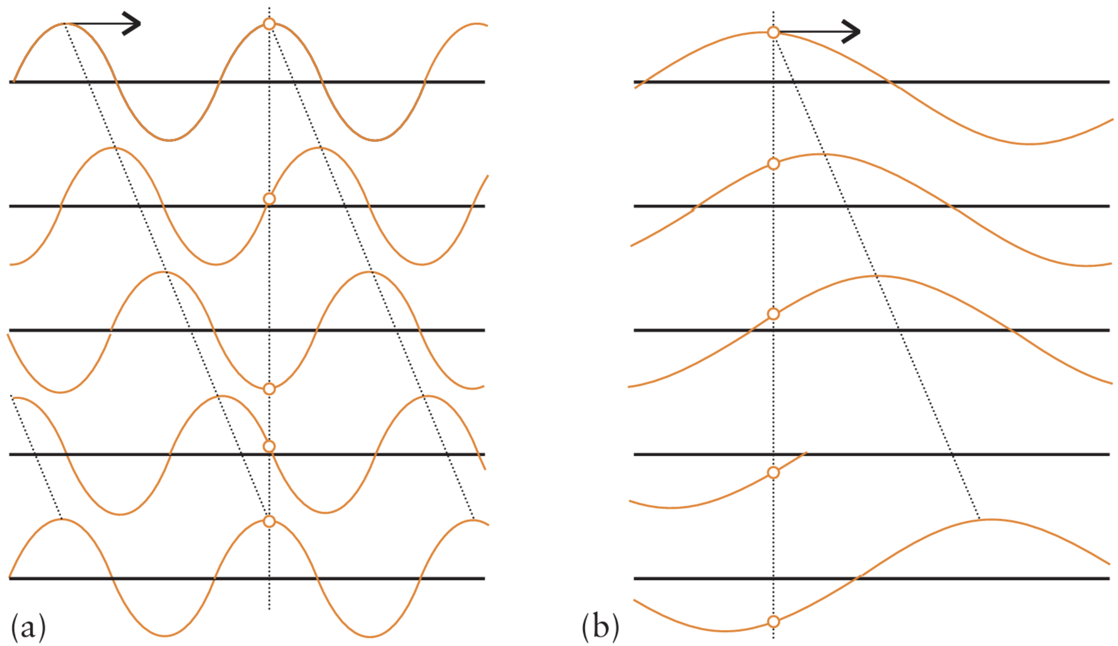


Figure 1A.8

1A.2 Electromagnetic radiation

Example 1A.1: **Calculating the wavelength of light of a known frequency**

Identify which color of light has the shorter wavelength: red light of frequency 4.3×10^{14} Hz or blue light of frequency 6.4×10^{14} Hz.

1A.2 Electromagnetic radiation

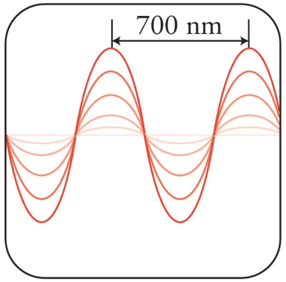
Example 1A.1: **Calculating the wavelength of light of a known frequency**

Identify which color of light has the shorter wavelength: red light of frequency 4.3 x 10^{14} Hz or blue light of frequency 6.4 x 10^{14} Hz.

SOLVE

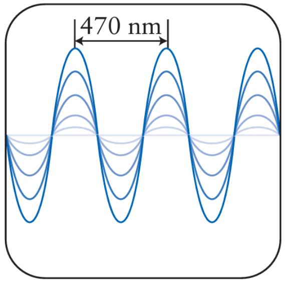
For red light: from $\lambda\nu = c$ written as $\lambda = c/\nu$,

$$\lambda = \frac{\overbrace{2.998 \times 10^{8}\ \text{m}\cdot\text{s}^{-1}}^{c}}{4.3 \times 10^{14}\ \underbrace{\text{s}^{-1}}_{\text{Hz}}} = \frac{2.998 \times 10^{8}}{4.3 \times 10^{14}}\ \text{m} = 7.0 \times 10^{-7}\ \text{m}$$



For blue light: from $\lambda\nu = c$ written as $\lambda = c/\nu$,

$$\lambda = \frac{\overbrace{2.998 \times 10^{8}\ \text{m}\cdot\text{s}^{-1}}^{c}}{6.4 \times 10^{14}\ \underbrace{\text{s}^{-1}}_{\text{Hz}}} = \frac{2.998 \times 10^{8}}{6.4 \times 10^{14}}\ \text{m} = 4.7 \times 10^{-7}\ \text{m}$$



1A.2 Electromagnetic radiation

The electromagnetic spectrum

TABLE 1.1 Color, Frequency, and Wavelength of Electromagnetic Radiation

Radiation type	Frequency (10^{14} Hz)	Wavelength (nm, 2 sf)*	Energy per photon (10^{-19} J)
x-rays and γ-rays	$\geq 10^3$	≤ 3	$\geq 10^3$
ultraviolet	8.6	350	5.7
visible light			
violet	7.1	420	4.7
blue	6.4	470	4.2
green	5.7	530	3.8
yellow	5.2	580	3.4
orange	4.8	620	3.2
red	4.3	700	2.8
infrared	3.0	1000	2.0
microwaves and radio waves	$\leq 10^{-3}$	$\geq 3 \times 10^6$	$\leq 10^{-3}$

*The abbreviation sf denotes the number of significant figures in the data. The frequencies, wavelengths, and energies are typical values; they should not be regarded as precise.

Visible light: 400-700 nm

White light: mixture of all wavelengths of visible light

Radiation of sun: white light plus shorter and longer wavelength radiation (UV and IR)

Figure 1A.9

1A.2 Electromagnetic radiation

The electromagnetic spectrum

- **Ultraviolet (UV) radiation**: shorter than violet, less than about 400 nm, causes sunburn and tanning, ozone layer protects us from it
- **Infrared (IR) radiation**: longer than red light, greater than about 800 nm, experienced as heat
- **Microwaves**: in the millimeter-to-centimeter range, used in radar and microwave ovens

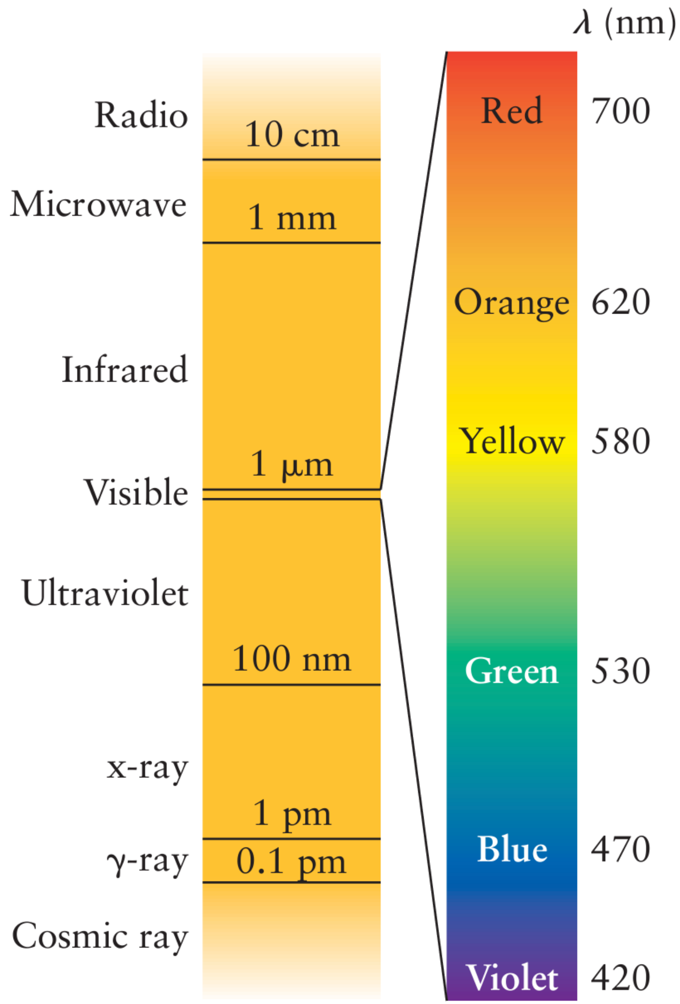
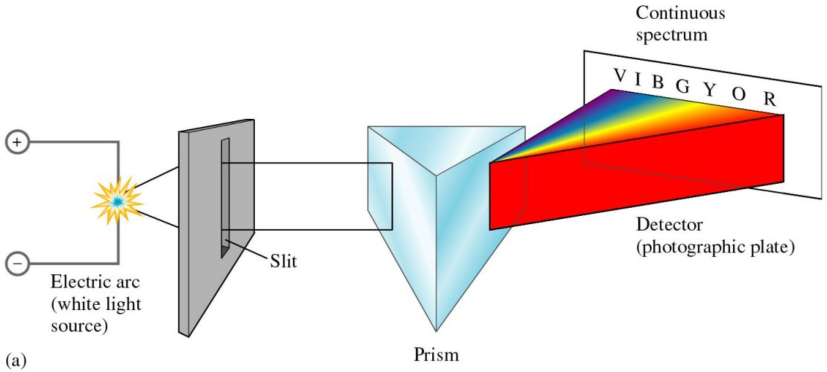


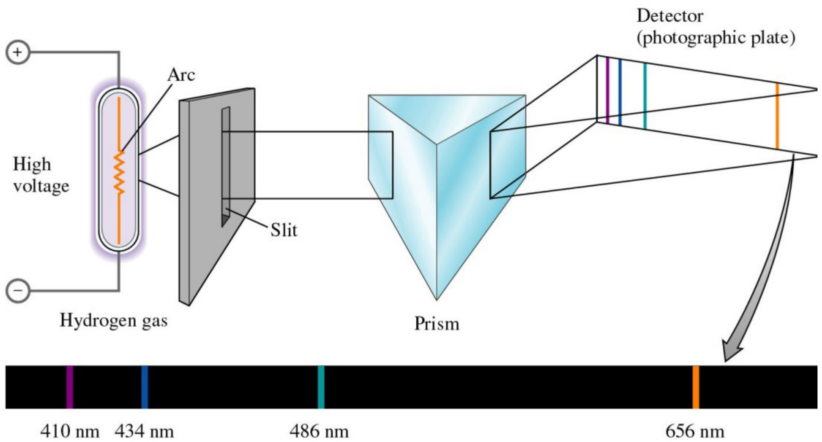
Figure 1A.9

1A.3 Atomic spectra

The atomic spectrum of hydrogen



- White light refracted through a prism → continuous spectrum (all visible wavelengths)



- Hydrogen gas + electric current → emits light
- Current excites H atoms → they release energy as electromagnetic radiation
- Through a prism → discrete **spectral lines** (not continuous)
- Bright red line at **656 nm**, plus lines in UV and IR

1A.3 Atomic spectra

The atomic (emission) spectrum of hydrogen

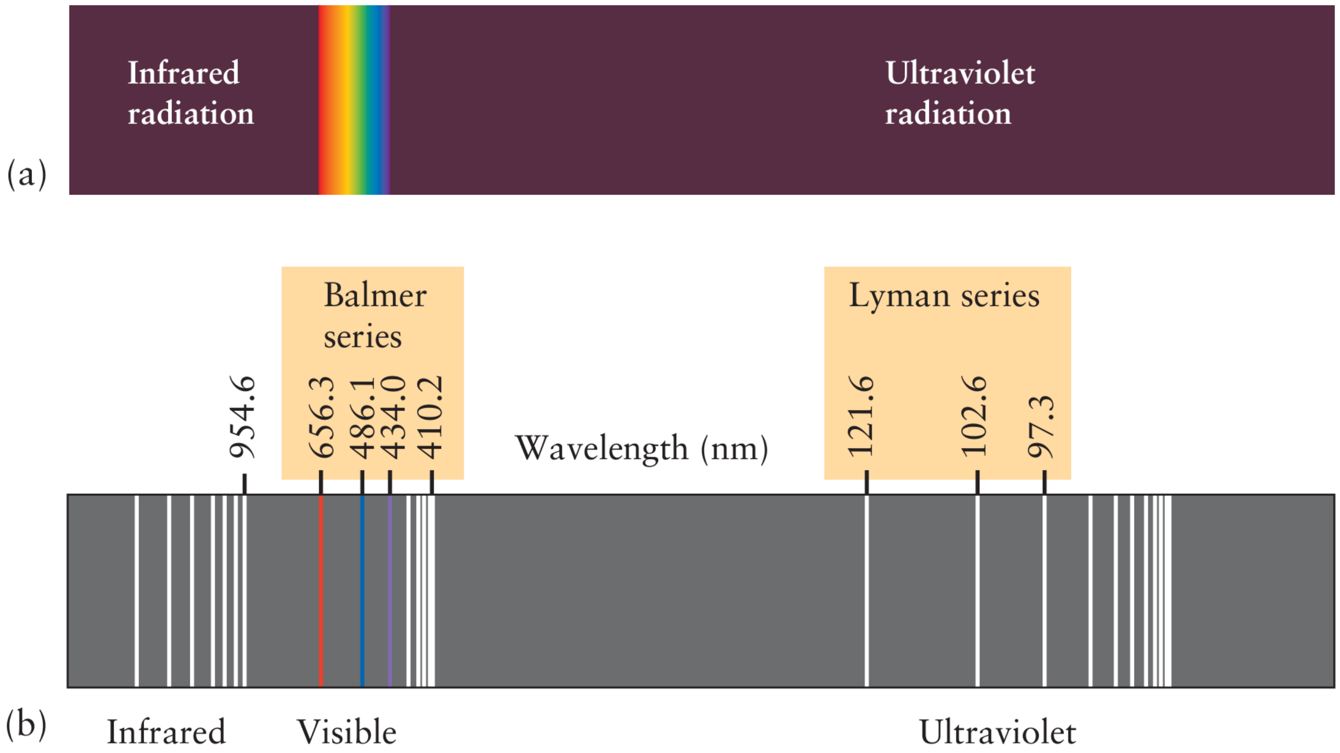


Figure 1A.10

(a) The infrared, visible, and ultraviolet spectrum.

(b) The complete **emission spectrum** of **atomic hydrogen**. The spectral lines have been assigned to various groups called series, two of which are shown with their names.

1A.3 Atomic spectra

The Rydberg formula

$$\nu = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \quad n_1 = 1,2,\ldots, n_2 = n_1 + 1, n_1 + 2, \ldots$$

Important formula.
Need to know how to apply.

With *R* is e empirical (experimentally determined) Rydberg constant; its value is 3.29 x 10^{15} Hz.

For now: With n_1 and n_2: positive integers, as shown above.

The Rydberg formula describes the wavelengths of the spectral lines in the **Balmer**, **Paschen**, and **Lyman series** for hydrogen.

The **Paschen** series is a set of lines in the **infrared** region with n_1 = 3 (and n_2 = 4, 5, . . .)
The **Balmer** series is the set of lines in the **visible** region with n_1 = 2 (and n_2 = 3, 4, . . .)
The **Lyman** series, a set of lines in the **ultraviolet** region of the spectrum, has n_1 = 1 (and n_2 = 2, 3, . . .).

1A.3 Atomic spectra

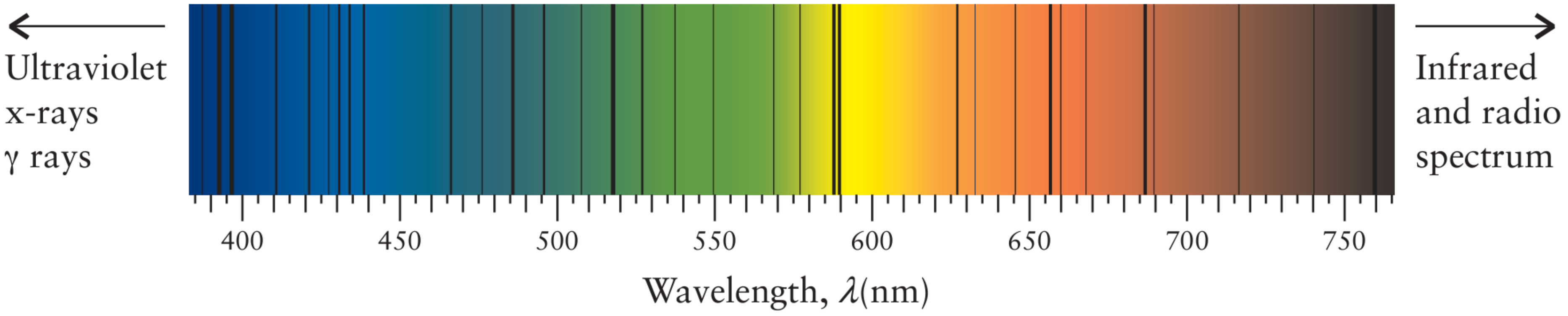
Example A1.2: **Identifying a line in the hydrogen spectrum**

Calculate the wavelength of the radiation emitted by a hydrogen atom for $n_1 = 2$ and $n_2 = 3$. Identify the spectral line in Fig. 1.10b.

1A.3 Atomic spectra

Absorption spectrum

Figure 1A.10



White light passed through a **vapor** composed of the **atoms of an element**
→ absorption spectrum: a series of dark lines on an otherwise continuous spectrum (Fig. 1.11).

- The absorption and emission lines have the same frequencies
- Suggests an atom can absorb and emitt radiation at same frequencies
- Absorption spectra are used by astronomers to identify elements in the outer layers of stars

The skills you have mastered are the ability to

- Describe the experiments that led to the formulation of the nuclear model of the atom
- Calculate the wavelength or frequency of light from the relation $\lambda\nu = c$ (example 1A.1)
- Calculate the wavelength of a transition in a hydrogen atom from the Rydberg formula (example 1A.2)

Summary: Scattering experiments show that an atom of element with atomic number *Z* consists of a tiny but massive nucleus surrounded by *Z* electrons. Electromagnetic radiation is a wave of characteristic frequency and wavelength that travels through space at speed *c*. Atomic spectroscopy is the analysis of the light emitted or absorbed by atoms. The observation of spectral lines strongly suggests that electrons in atoms can have only certain energies.