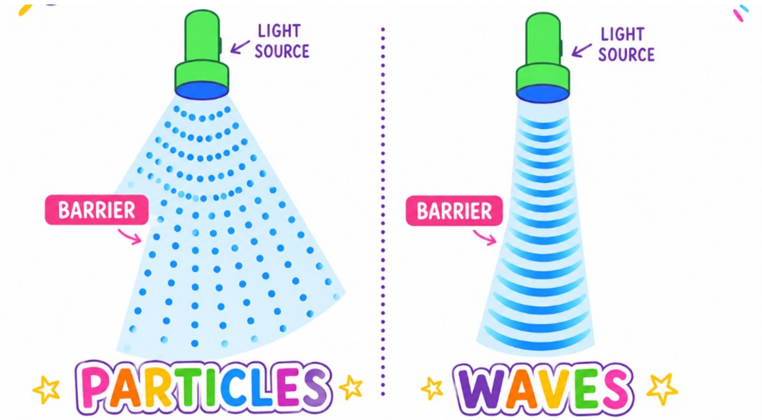
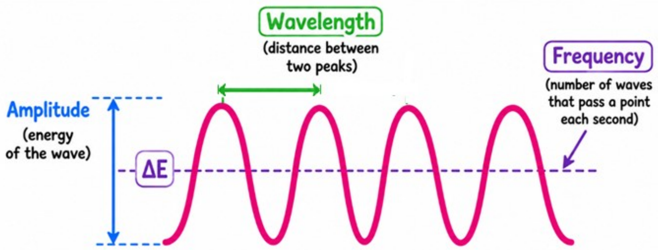
EM radiation

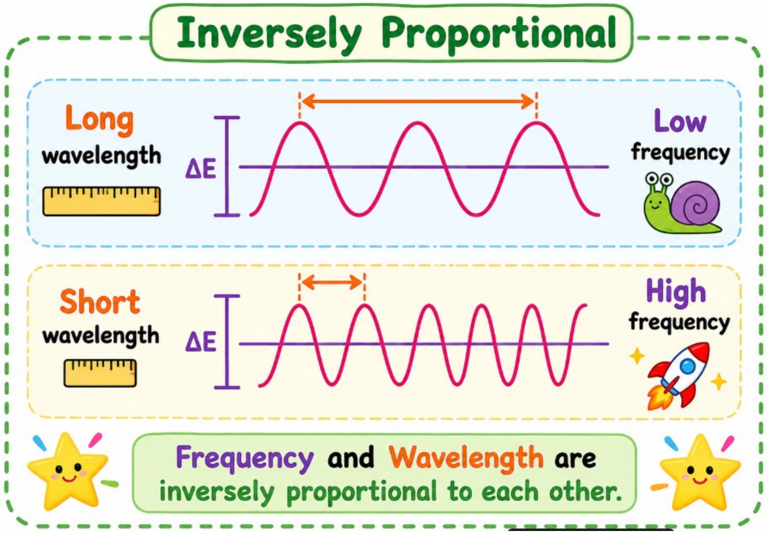
EM radiation can be both a wave/particle.



Wave model – EM Radiation

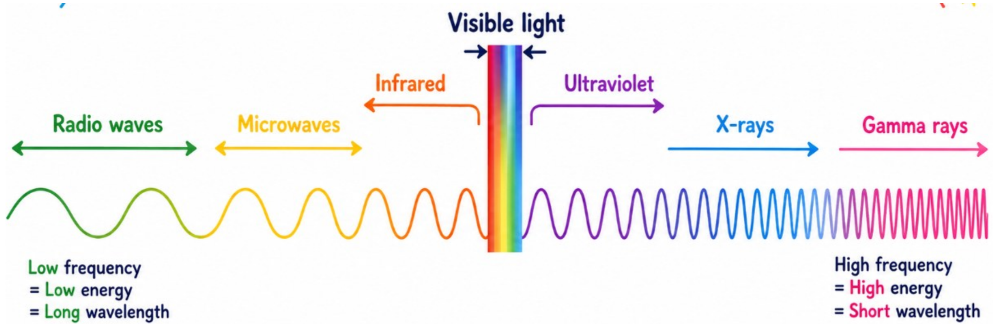


1. **frequency** (f)
   - Number of waves that pass a fixed point in 1 second.
   (Hz or s^{-1})
2. **wavelength** (λ)
   - Distance between two crests or troughs.
   (nm = 10^{-9} m)



Electromagnetic Spectrum

Described in terms of waves and characterised in terms of wavelength and frequency.



Energy is **PROPORTIONAL TO** the **frequency** of radiation!

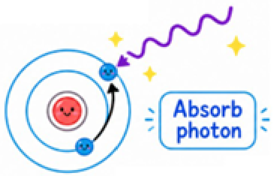
high frequency (low wavelength) ⟷ high energy

Photon

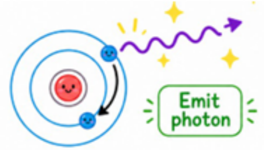
EM radiation can behave like a stream of particles when absorbed/emitted from matter.

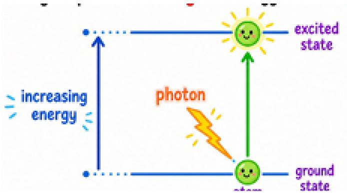
These photons transfer small bundles (packets) of energy.

Energy is gained by electrons within the substance when a photon is absorbed by matter.



Energy is lost by electrons within the substance when a photon is emitted by matter.





The energy of a mole of photons is calculated using:

For 1 mole of photons $E = \dfrac{L\,h\,c}{1000\,\lambda}$

The unit of E is kJ mol^{-1}

E	L	h	c	λ (lambda)	1000
Energy	Avogadro constant	Planck's constant	Speed of light	Wavelength	Conversion factor
This is the energy of 1 mole of photons.	The number of particles in 1 mole.	A very small constant that relates energy and frequency.	The speed of light in a vacuum.	The distance between successive peaks of a wave.	Dividing by 1000 converts the energy from joules (J) to kilojoules (kJ).
Unit: kJ mol^{-1}	Value: 6.02×10^{23} Unit: mol^{-1}	Value: 6.63×10^{-34} Unit: J s	Value: 3.0×10^{8} Unit: m s^{-1}	Unit: m. If λ is given in nm, multiply by 10^{-9} to convert to m.	Unit: none

1 MOLE OF PHOTONS

(energy in kJ mol^{-1})

Worked Example

Camphorquinone generates free radicals when exposed to radiation with a wavelength of 471 nm.

Calculate the energy, in kJ mol^{-1}, associated with this wavelength.

Constants used:
$L = 6.02 \times 10^{23}\ mol^{-1}$
$h = 6.63 \times 10^{-34}$ J s
$c = 3.00 \times 10^{8}\ m\ s^{-1}$

$$E = \frac{L\,h\,c}{1000\,\lambda}$$

Energy unit = kJ mol^{-1}

500 nm = 500×10^{-9} m (start with nanometres; $\times 10^{-9}$ to convert to metres)

Step 1 Use $E = \dfrac{Lhc}{\lambda}$ to calculate energy per mole.

$$E = 6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 3 \times 10^{8} \times \frac{1}{471 \times 10^{-9}}$$

(L) Avogadro constant (mol^{-1}); (h) Planck's constant (J s); (c) Speed of light (m s^{-1}); (λ) Wavelength (m)

Step 2 Calculate the value.

$$\lambda = \frac{0.1197}{471 \times 10^{-9}} = 2.54 \times 10^{5}\ J\ mol^{-1} = 254\ kJ\ mol^{-1}$$

Energy = 254 kJ mol^{-1}

Worked Example

The free radical benzoyl peroxide absorbs radiation with a wavelength of 365 nm to break the O–O bond.

Calculate the energy, in kJ mol^{-1}, associated with this radiation.

Constants used:
$L = 6.02 \times 10^{23}\ mol^{-1}$
$h = 6.63 \times 10^{-34}$ J s
$c = 3.00 \times 10^{8}\ m\ s^{-1}$

$$E = \frac{L\,h\,c}{1000\,\lambda}$$

Energy unit = kJ mol^{-1}

365 nm = 365×10^{-9} m (start with nanometres; $\times 10^{-9}$ to convert to metres)

Step 1 Use $E = \dfrac{Lhc}{\lambda}$ to calculate energy per mole.

$$E = 6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 3 \times 10^{8} \times \frac{1}{365 \times 10^{-9}}$$

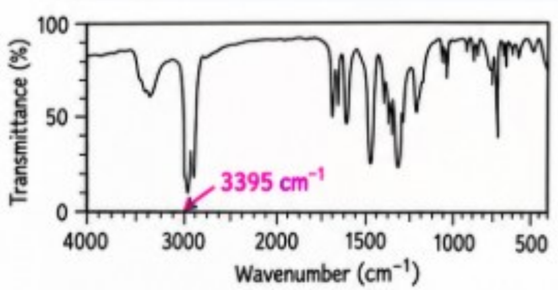
(L) Avogadro constant (mol^{-1}); (h) Planck's constant (J s); (c) Speed of light (m s^{-1}); (λ) Wavelength (m)

Step 2 Calculate the value.

$$\lambda = \frac{0.1197}{365 \times 10^{-9}} = 3.28 \times 10^{5}\ J\ mol^{-1} = 328\ kJ\ mol^{-1}$$

Energy = 328 kJ mol^{-1}

WAVENUMBER TO WAVELENGTH



Formula:

$$\text{wavelength (m)} = \frac{1}{\text{wavenumber} \div 100}$$

IR

The infrared spectrum of altizarin shows a peak at 3395 cm^{-1}.
Calculate (i) the **wavelength**, in metres
and (ii) the **energy**, in $kJ\ mol^{-1}$, of this radiation.

Step 1 Convert wavenumber to wavelength (in metres).

Formula:

$$\lambda = \frac{1}{\tilde{\nu}}$$

Substitute the values:

× 100 to convert to m^{-1}

$$\lambda = \frac{1}{3395\ cm^{-1}} \quad \lambda = \frac{1}{3395 \times 100\ m^{-1}}$$

Calculate:

$$\lambda = 2.95 \times 10^{-6}\ m$$

Wavelength = 2.95×10^{-6} m

Step 2 Calculate the energy (in $kJ\ mol^{-1}$).

Formula:

$$E = \frac{L\,h\,c}{\lambda} = 6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 3.00 \times 10^{8} \times \frac{1}{2.95 \times 10^{-6}}$$

(L) Avogadro constant (mol^{-1}); (h) Planck's constant (J s); (c) Speed of light ($m\ s^{-1}$); (λ) Wavelength (m)

Calculate:

$$E = \frac{(6.02 \times 10^{23})(6.63 \times 10^{-34})(3.00 \times 10^{8})}{2.95 \times 10^{-6}} = 4.06 \times 10^{4}\ J\ mol^{-1} = 40.6\ kJ\ mol^{-1}$$

Energy = 40.6 $kJ\ mol^{-1}$

Worked Example

A firework produced coloured light with energy of 193 $kJ\ mol^{-1}$.
Calculate the **wavelength**, in **nm**, of this coloured light.

Use the formula:

$$E = \frac{L\,h\,c}{1000\,\lambda}$$

E = energy ($kJ\ mol^{-1}$)
L = Avogadro constant ($6.022 \times 10^{23}\ mol^{-1}$)
h = Planck's constant (6.626×10^{-34} J s)
c = speed of light ($2.998 \times 10^{8}\ m\ s^{-1}$)
λ = wavelength (m)

Remember to change energy value into joules (×1000)

Rearrange the formula to make λ the subject:

$$\lambda = \frac{L\,h\,c}{1000\,E}$$

Substitute the values:

$$\lambda = \frac{(6.022 \times 10^{23})(6.626 \times 10^{-34})(2.998 \times 10^{8})}{1000 \times 193}$$

Calculate the numerator (top):

$(6.022 \times 10^{23})(6.626 \times 10^{-34})(2.998 \times 10^{8}) = 1.197 \times 10^{-1}$

Now divide by the denominator (bottom):

$$\lambda = \frac{1.197 \times 10^{-1}}{193000} = 6.22 \times 10^{-7}\ m$$

622 nm

A sparkler produced coloured light with energy of 247 $kJ\ mol^{-1}$.
Calculate the **wavelength**, in **nm**, of this coloured light.

Use the formula:

$$E = \frac{L\,h\,c}{1000\,\lambda}$$

E = energy ($kJ\ mol^{-1}$)
L = Avogadro constant ($6.022 \times 10^{23}\ mol^{-1}$)
h = Planck's constant (6.626×10^{-34} J s)
c = speed of light ($2.998 \times 10^{8}\ m\ s^{-1}$)
λ = wavelength (m)

Remember to change energy value into joules (×1000)

Rearrange the formula to make λ the subject:

$$\lambda = \frac{L\,h\,c}{1000\,E}$$

Substitute the values:

$$\lambda = \frac{(6.022 \times 10^{23})(6.626 \times 10^{-34})(2.998 \times 10^{8})}{1000 \times 247}$$

Calculate the numerator (top):

$(6.022 \times 10^{23})(6.626 \times 10^{-34})(2.998 \times 10^{8}) = 1.195 \times 10^{-1}$

Now divide by the denominator (bottom):

$$\lambda = \frac{1.195 \times 10^{-1}}{247000} = 4.84 \times 10^{-7}\ m$$

484 nm

FOR 1 PHOTON

(energy in J)

USING WAVELENGTH (λ)

$$E = \frac{hc}{\lambda}$$

This gives energy in J (no need to divide by 1000)

Worked Example

A certain atom absorbs radiation with a wavelength of 620 nm.
Calculate the energy, in joules (J), of a single photon of this radiation.

Constants used:
$h = 6.63 \times 10^{-34}$ J s
$c = 3.00 \times 10^{8}$ m s^{-1}

Remember to change wavelength into metres (m) by 10^{-9}

Energy of 1 photon = J

$620 \text{ nm} = 620 \times 10^{-9}$ m

start with nanometres

$\times 10^{-9}$ to convert to metres

Step 1 Use $E = \frac{hc}{\lambda}$ to calculate energy of one photon.

$$E = 6.63 \times 10^{-34} \times 3.00 \times 10^{8} \div 620 \times 10^{-9}$$

(h) Planck's constant (J s)

(c) Speed of light (m s^{-1})

(λ) Wavelength (m)

Step 2 Calculate the value.

$$E = \frac{6.63 \times 10^{-34} \times 3.00 \times 10^{8}}{620 \times 10^{-9}}$$

$$= 3.20 \times 10^{-19} \text{ J}$$

Energy of 1 photon = 3.20×10^{-19} J

FOR 1 PHOTON

(energy in J)

USING FREQUENCY (f)

$$E = hf$$

This gives energy in J (no need to divide by 1000)

Worked Example

A sample of strontium was exposed to electromagnetic radiation with a frequency of $3{\cdot}08 \times 10^{17}$ s^{-1}.
Calculate the energy, in J, of this electromagnetic radiation.

Constants used:
$h = 6.63 \times 10^{-34}$ J s
$f = 3.08 \times 10^{17}$ s^{-1}

Use the formula:
$E = h \times f$

E = energy (J)
h = Planck's constant (J s)
f = frequency (s^{-1})

Step 1 Substitute the values into the formula.

$$E = (6.63 \times 10^{-34} \text{ J s}) \times (3.08 \times 10^{17} \text{ s}^{-1})$$

Step 2 Calculate the value.

$$E = 6.63 \times 10^{-34} \times 3.08 \times 10^{17}$$

$$E = 2.04 \times 10^{-16} \text{ J}$$

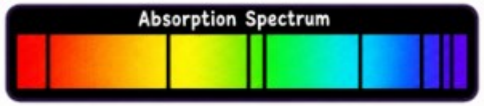
Energy of this electromagnetic radiation = 2.04×10^{-16} J

Spectroscopy

This shows a series of lines at discrete energy levels as photons are absorbed/emitted from atoms.

Absorption spectrum

measuring how the intensity of absorbed light varies with wavelength.

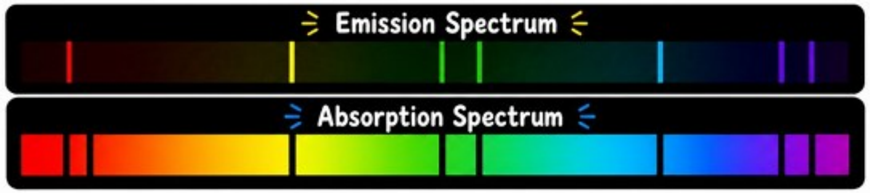


Emission spectrum

Measuring how the intensity of light emitted varies with wavelengths

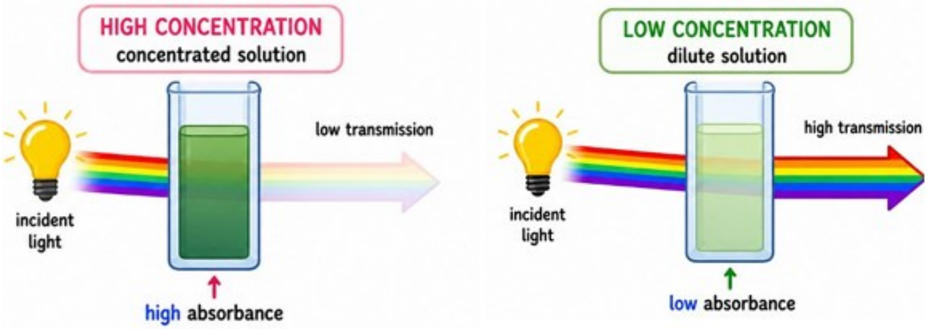
Each element in a sample produces characteristic absorption/emission spectra.

1. Identify element(s) in sample
2. Quantify/Concentration element in sample



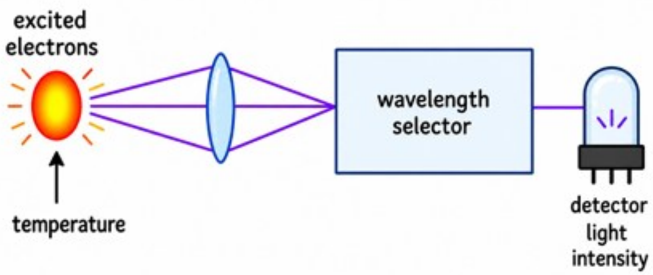
Absorption Spectra

Higher the concentration = Higher absorbance OR Lower transmission



Emission Spectra

High temperatures are used to excite the electrons within atoms.



Lines are formed by electrons releasing energy as they fall to lower energy levels.