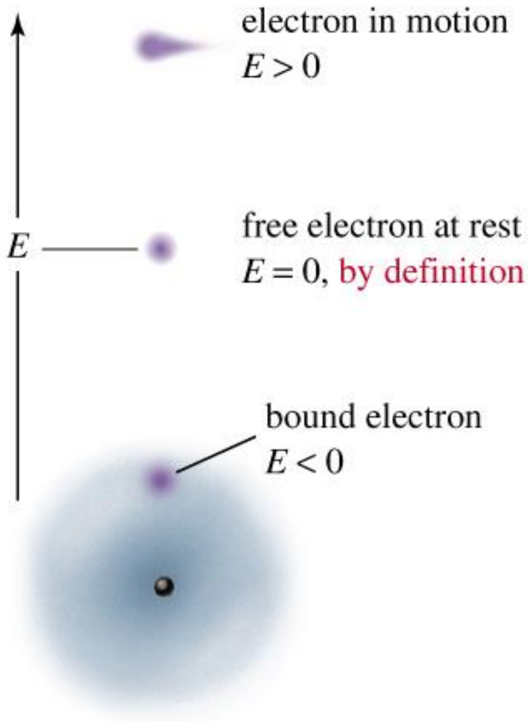
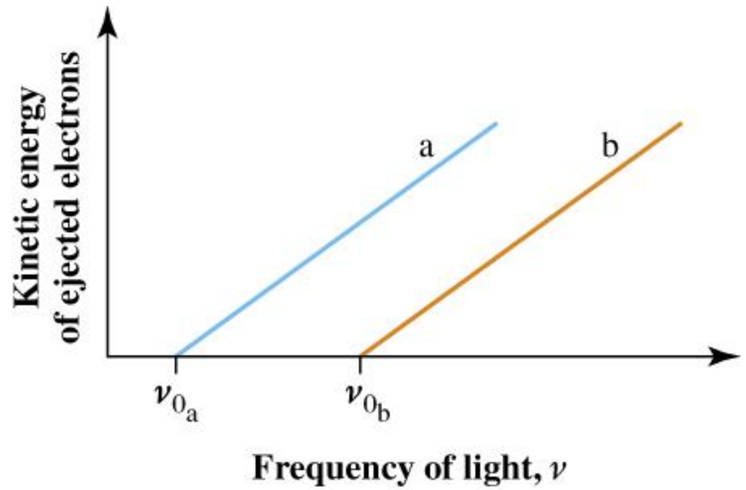
CHEM1011 LECTURE 8

Dr Shannan Maisey

ELECTRON ENERGIES





Free, moving electrons have kinetic energy. The energies these electrons can have are continuous, as seen from the photoelectric effect experiment.

Classical mechanics

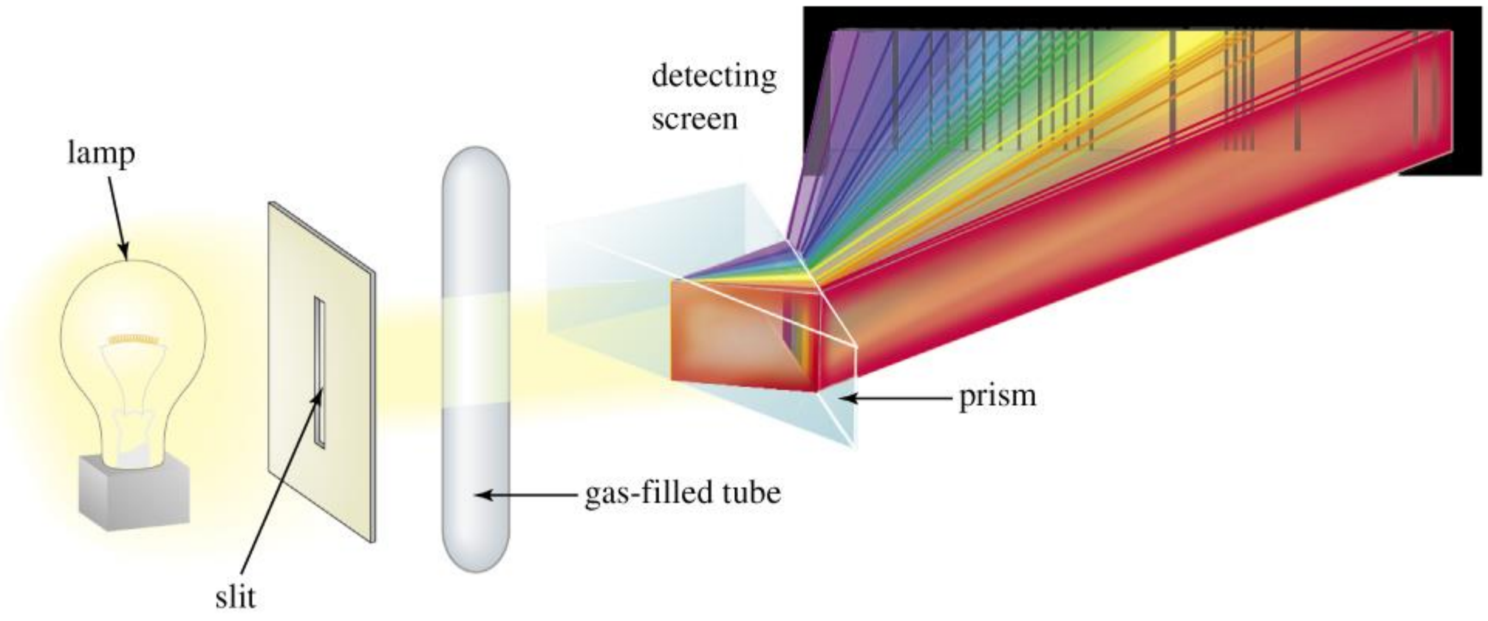
What about the energies of electrons that are bound in atoms?

Why was a 'threshold' frequency required in the photoelectric experiment?

Quantum mechanics

INTERACTION OF LIGHT AND MATTER

White light has photons with a range of wavelengths (polychromatic). When they pass through the gaseous atoms some of the photons' energies (i.e. specific frequencies) *exactly match* the difference between energy levels of an atom. These photons are consequently **absorbed,** leading to no light of that frequency/wavelength showing on the detector (i.e. the black lines).



QUANTISATION

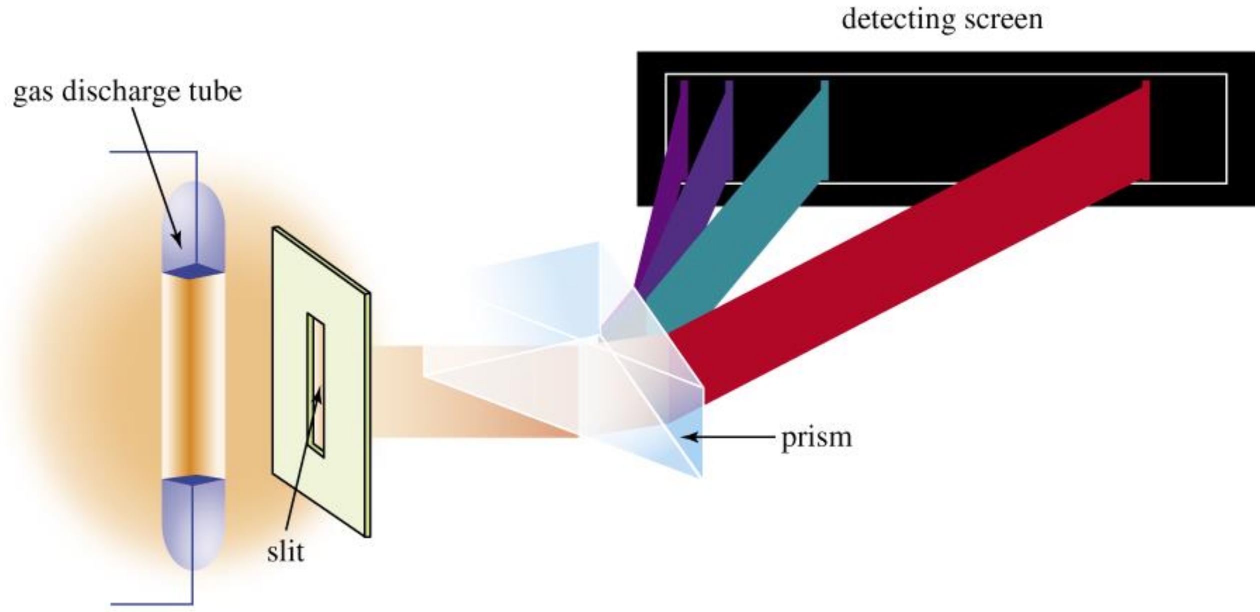
This pattern is known as an absorption spectrum; **light of specific frequencies has been absorbed by atoms**, causing the black lines. But being below the threshold frequency: the electron is not lost from the atom….

Specific frequencies/energies are being absorbed by the atom. This indicates **the energies available in an atom are non-continuous**. Rather, there are **discrete** 'steps' of energy available, like rungs on a ladder. The energy is restricted to certain levels, known as **quantisation**.

When an atom absorbs energy from photons, its only those photons with complimentary *frequency/energy* to the energy levels the atom contains (taking a step).

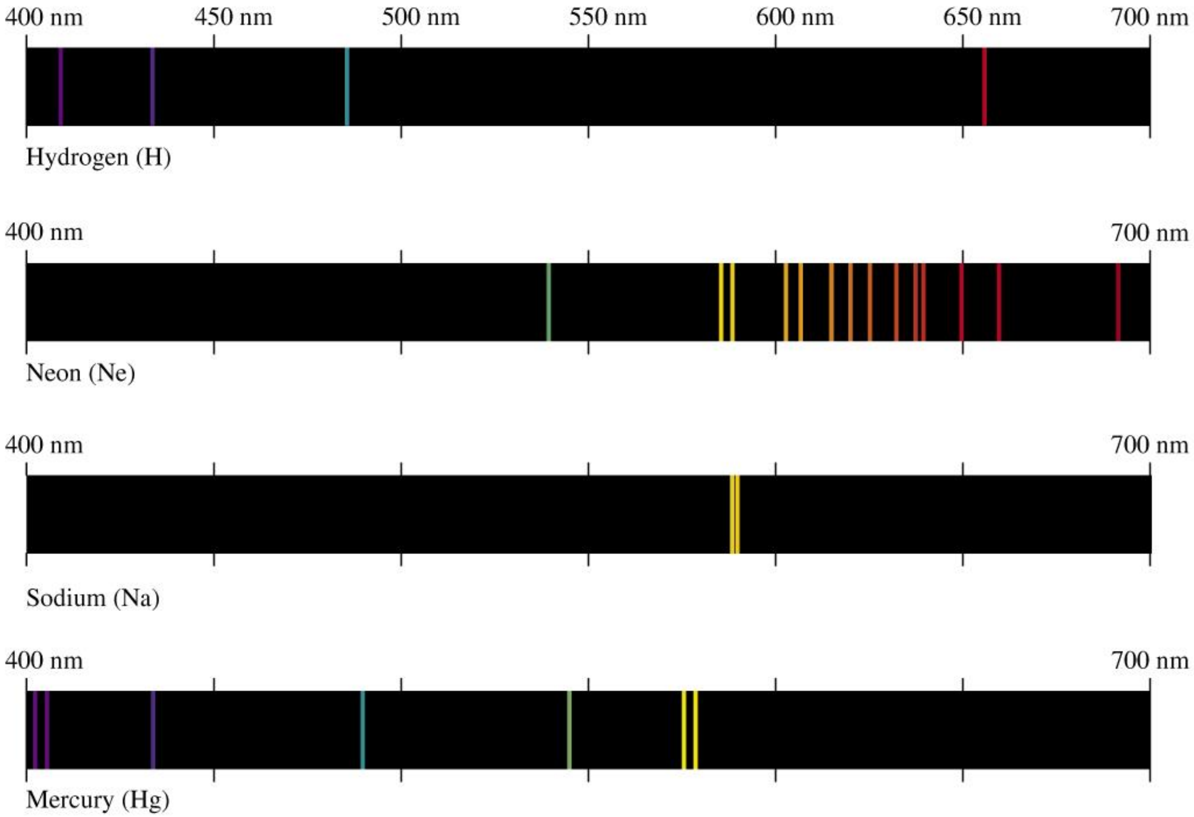
EMISSION

A similar experiment where gases were heated by an electrical discharge (current passed through) produces light at specific frequencies (lines). Since light (photons) was **emitted** by the gaseous atoms the spectrum produced is known as an **emission spectrum.**



EMISSION

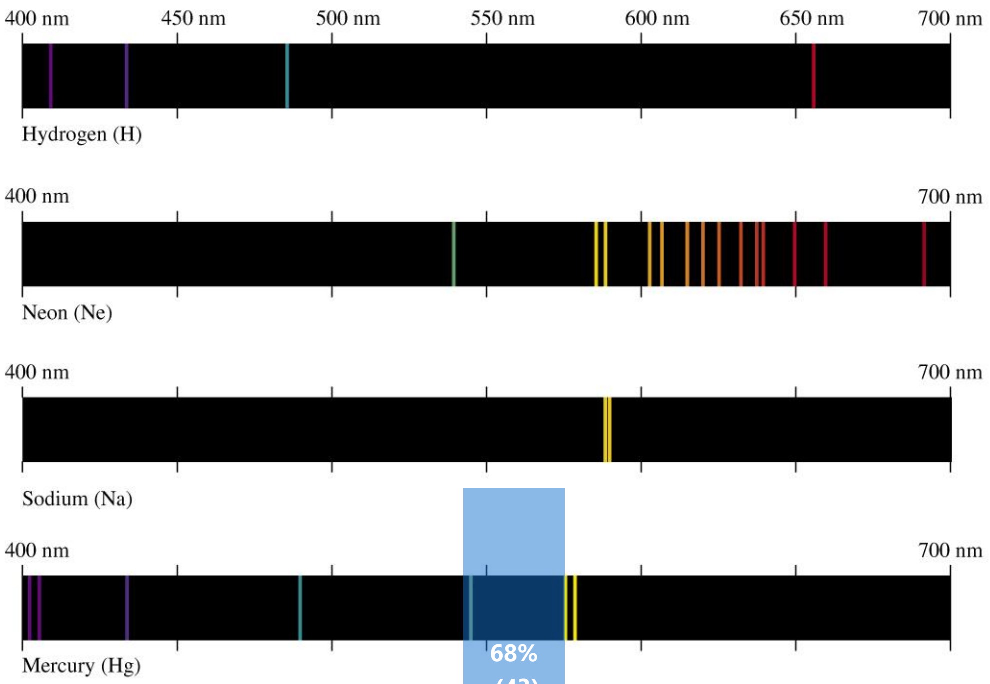
Like absorption, the emission lines correspond to *exact energy level differences* in the gaseous atoms however now they are due to a photon of specific frequency/energy being emitted instead of absorbed. It’s the exact opposite of absorption.

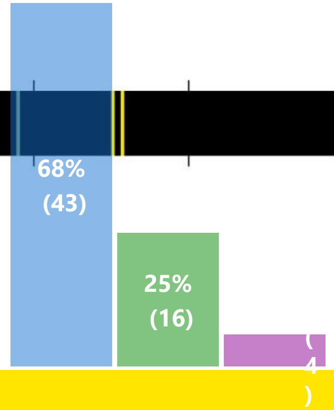


THINK CRITICALLY!

Are the lines in an absorption spectrum for a certain element **at the same wavelength** for the emission spectrum of that element? (Do the lines match)

- A Yes ✓
- B No
- C Depends on the element
- D Don't know





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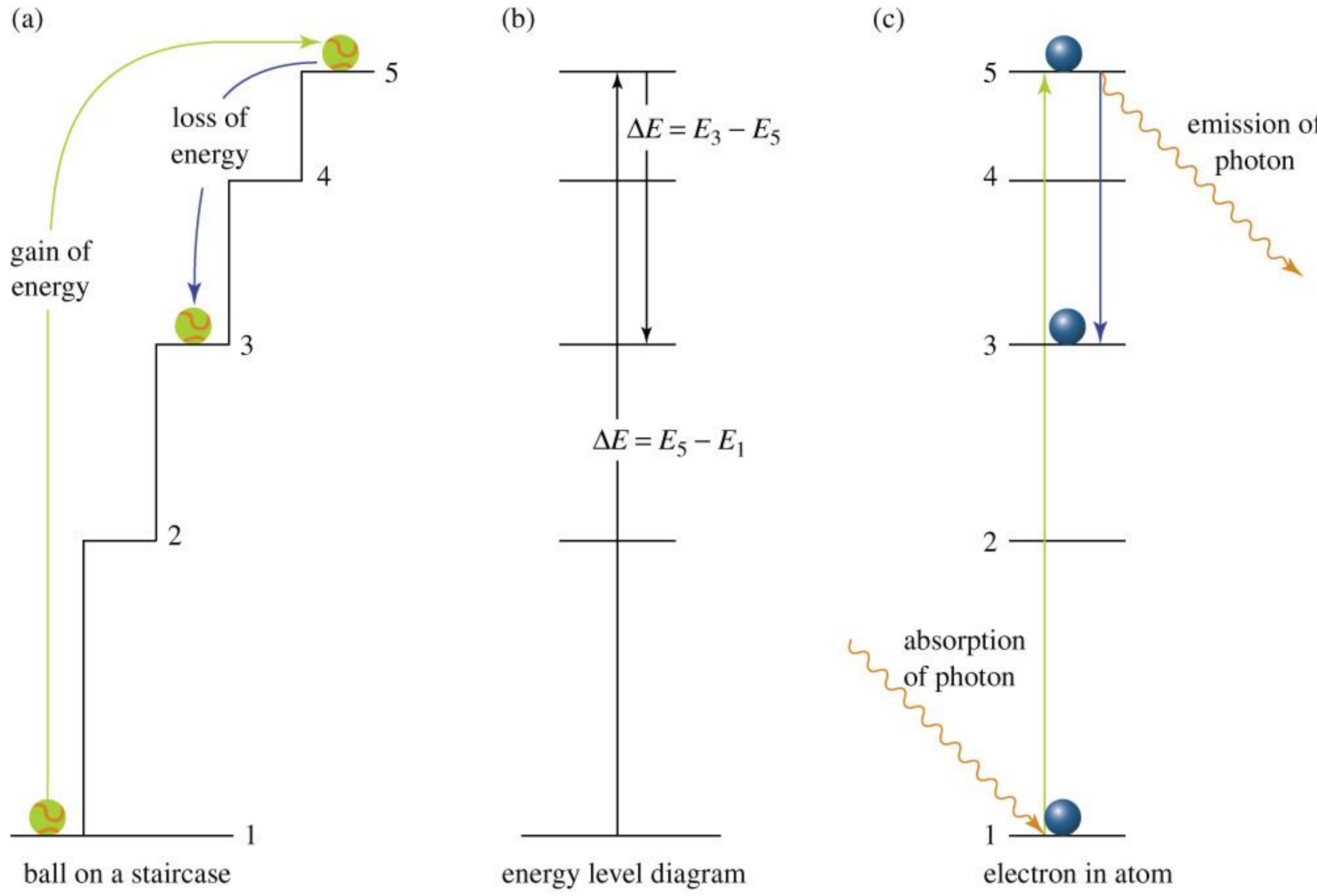
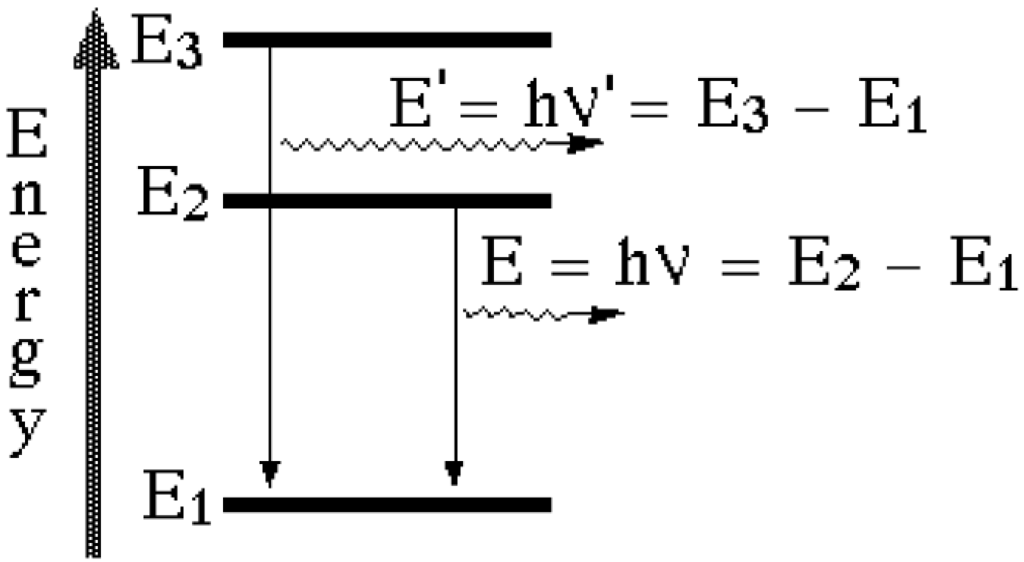
ENERGY LEVEL DIAGRAM



QUANTISATION OF ENERGY

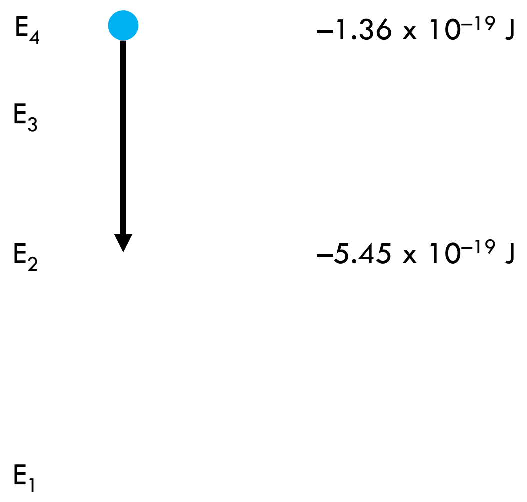
When an electron goes from a higher energy level to a lower level, a photon is emitted with energy, **E = hν**, equal to the <u>difference</u> between the two levels.



WORKED EXAMPLE 4.4

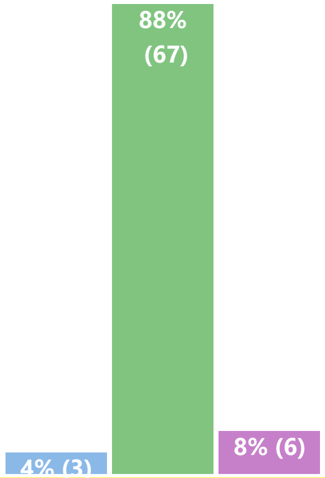
What is the energy change when the electron in a hydrogen atom undergoes a transition from the fourth energy level to the second energy level?

Use $E_4 = -1.36 \times 10^{-19}$ *J and* $E_2 = -5.45 \times 10^{-19}$ *J*



Q: This transition corresponds to:

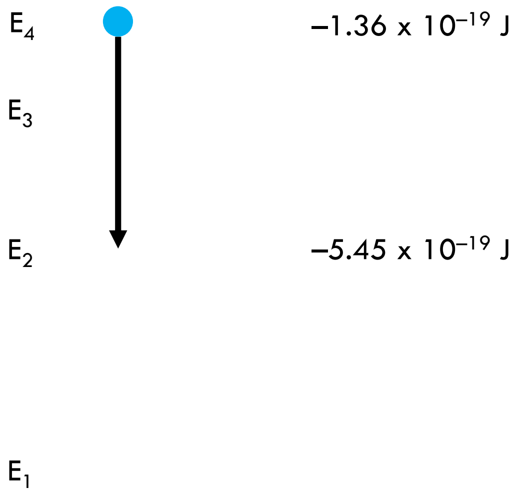
- **A** Absorption
- **B** Emission
- **C** Explosion
- **D** Don't know



WORKED EXAMPLE 4.4

What is the energy change when the electron in a hydrogen atom undergoes a transition from the fourth energy level to the second energy level? What is the wavelength of the photon emitted?

Use $E_4 = -1.36 \times 10^{-19}$ J *and* $E_2 = -5.45 \times 10^{-19}$ J



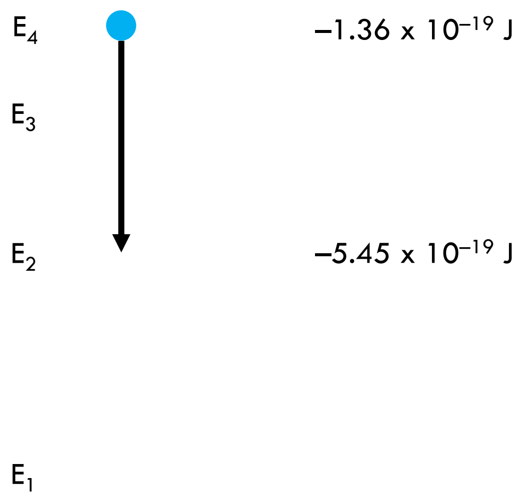
Final - initial

$$\Delta E = E_2 - E_4$$
$$= -5.45 \times 10^{-19} - (-1.36 \times 10^{-19})$$
$$= -4.09 \times 10^{-19} J$$

WORKED EXAMPLE 4.4

What is the energy change when the electron in a hydrogen atom undergoes a transition from the fourth energy level to the second energy level? **What is the wavelength of the photon emitted?**

Use $E_4 = -1.36 \times 10^{-19}$ J and $E_2 = -5.45 \times 10^{-19}$ J



$$E_{photon} = h\nu = hc/\lambda$$

$$\lambda_{photon} = \frac{hc}{E_{photon}}$$

$$= \frac{(6.626 \times 10^{-34} Js) \times (2.998 \times 10^{8} ms^{-1})}{4.09 \times 10^{-19} J}$$

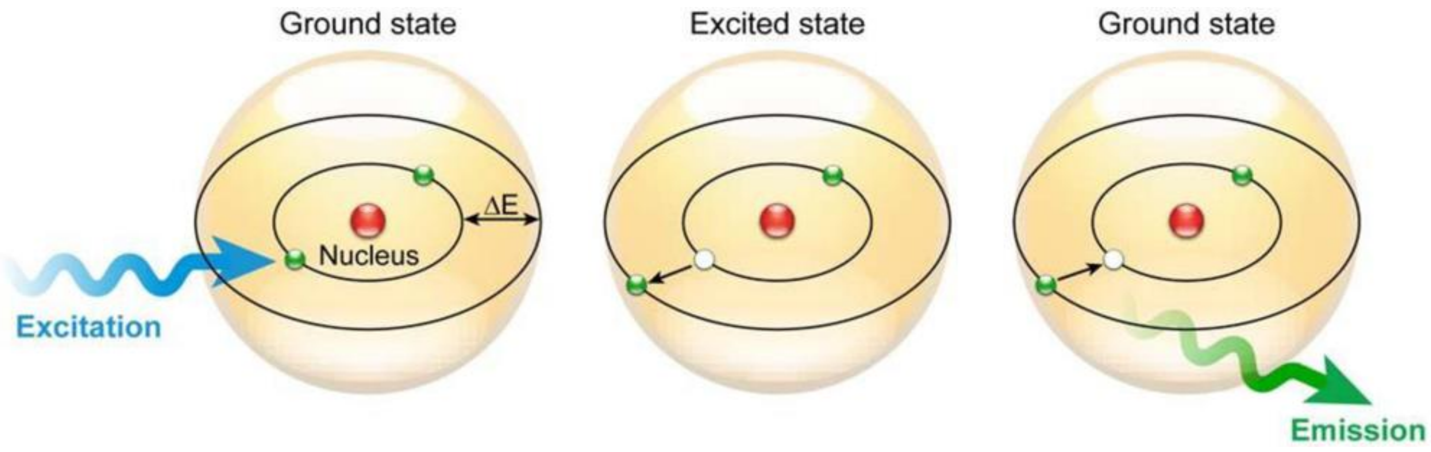
$$= 486 \times 10^{-9} m$$

Or 486 nm

GROUND AND EXCITED STATES

Ground state of an atom is one in which it has the **lowest (most stable) energy**.

Excited states are any energy level higher than the ground state (unstable). Excited atoms subsequently give up their excess energy (i.e. emit photons or in collisions) to return to a lower energy state (eventually ending back at the ground state).



THE AURORA-AUSTRALIS

Electrons in N_2 and O_2 gas molecules are moved to excited states after being hit with high energy from the sun.

As the electrons transition to their ground state they emit photons of light, mainly in wavelengths corresponding to green and purple.

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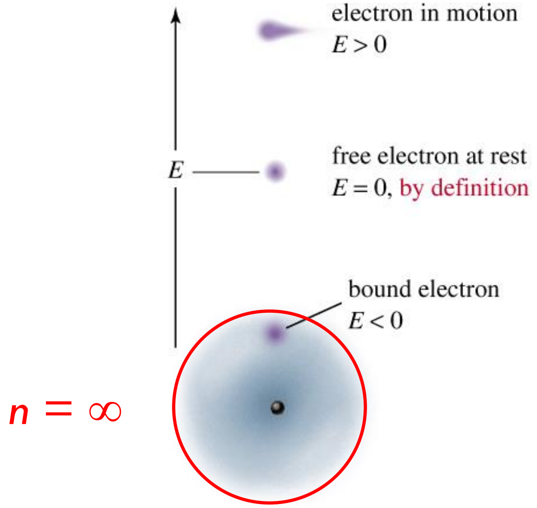
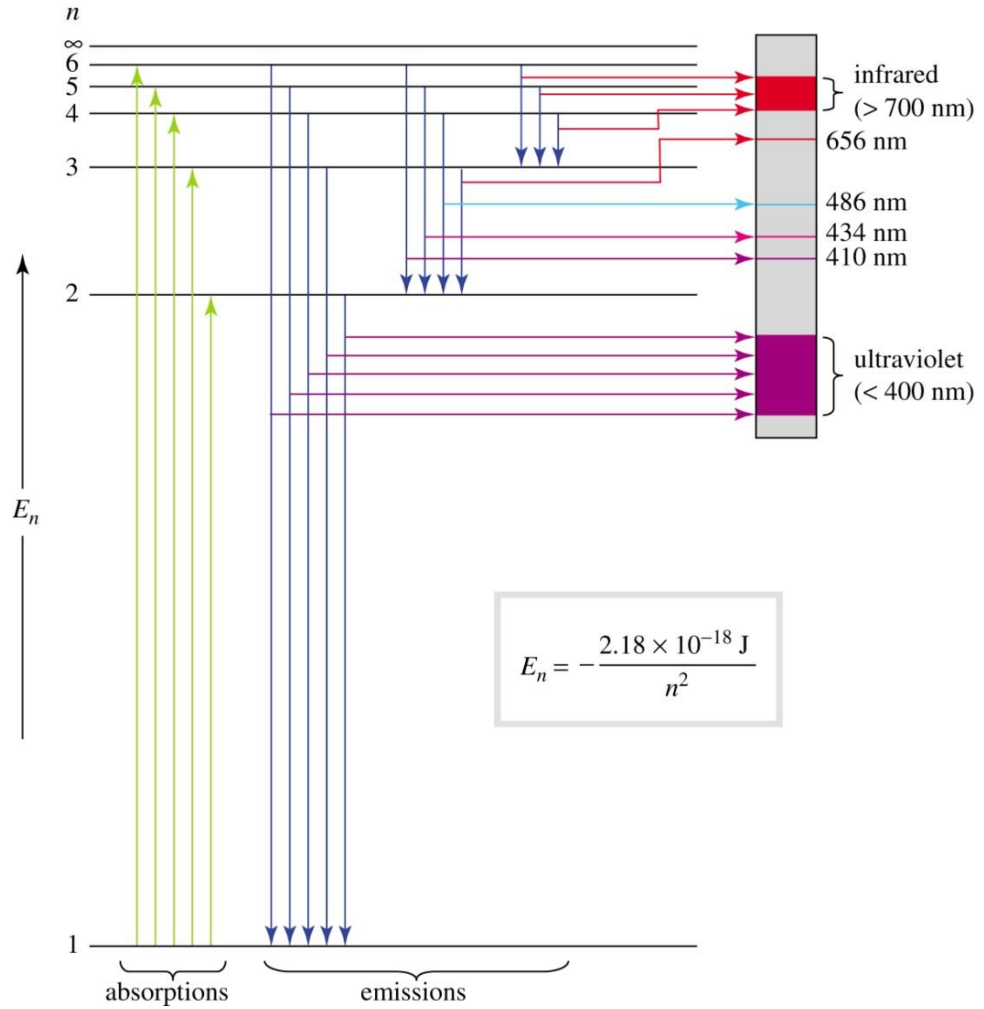
HYDROGEN'S ATOMIC ENERGY LEVELS

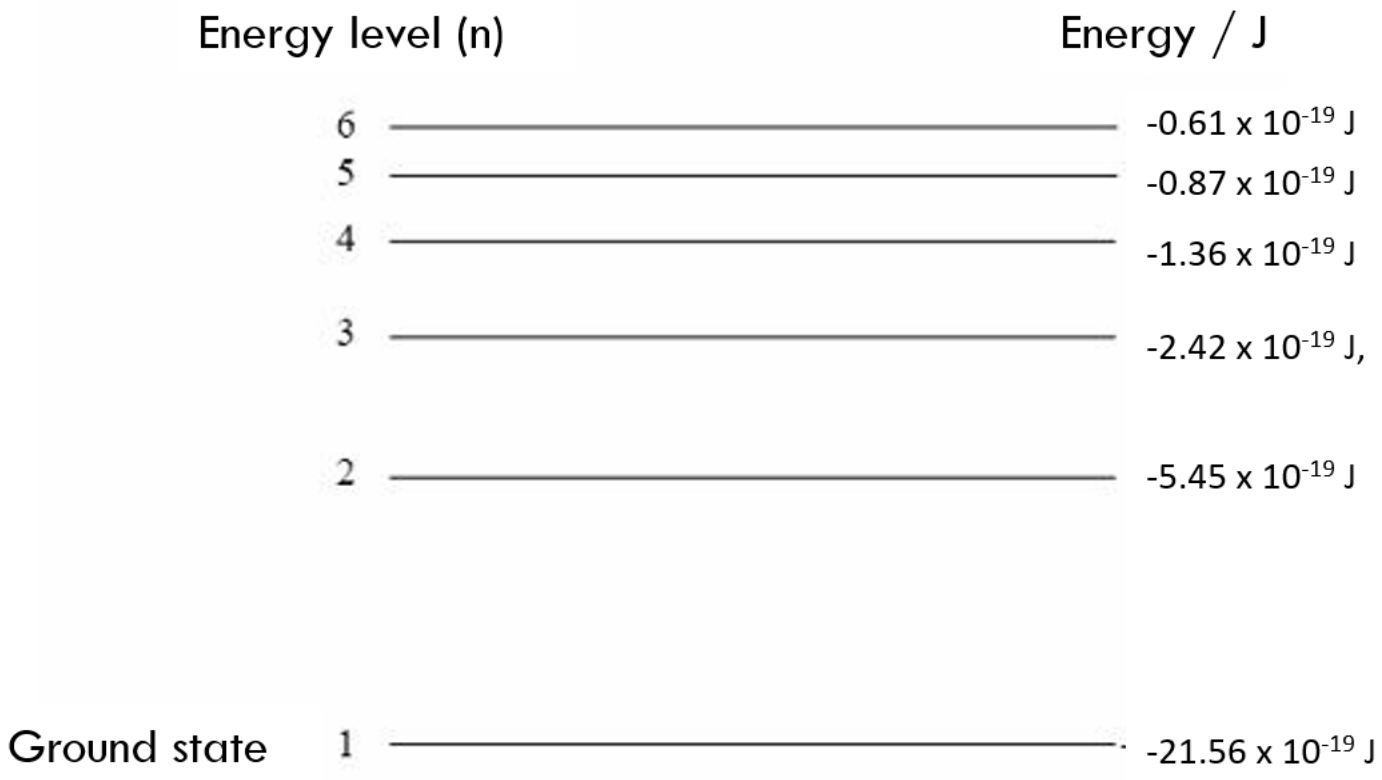
A hydrogen atom has a **regular progression of quantised energy levels** (*other atoms do too but it gets complicated with more electrons*).

Energy levels become increasingly closer in energy (bunch together) the further they are from the nucleus. Beyond these energy levels the electron is no longer bound by the nucleus and the atom has been **ionised**.





YOU TRY: RAW ALL OF THE (EMISSION) ENERGY TRANSITIONS IN A H ATOM



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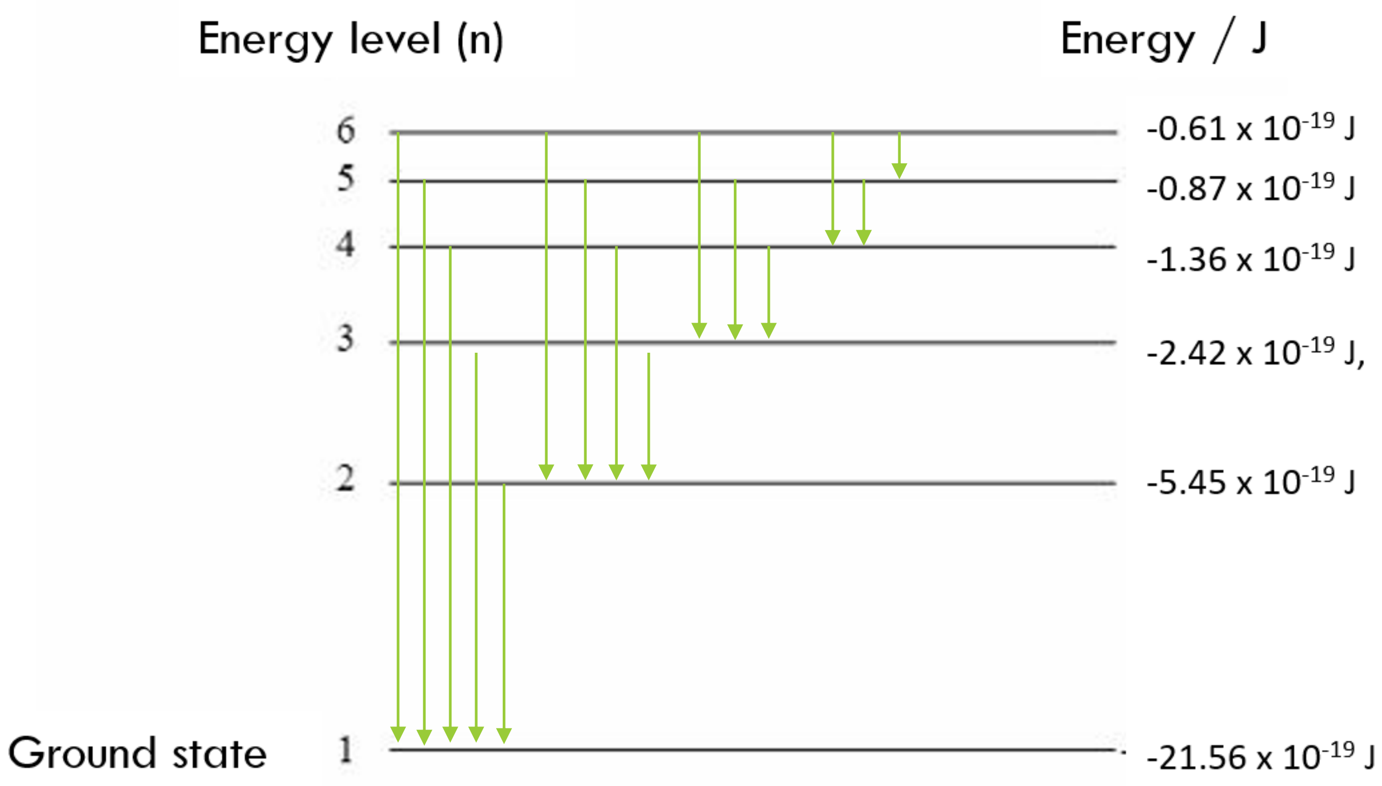


BALMER SERIES

Swiss mathematician and high school teacher **Johann Balmer** discovered in 1885 a mathematical relationship which was able to describe each of the emissions of the hydrogen atom.

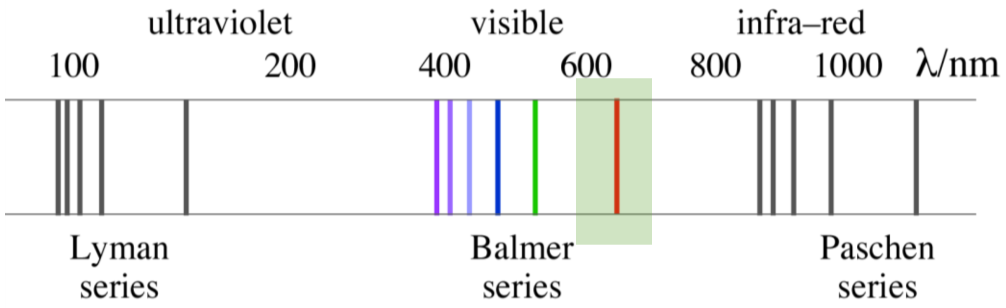
https://en.wikipedia.org/wiki/Johann_Jakob_Balmer

The visible parts of the hydrogen atomic spectrum in the Balmer series (consisting of the wavelengths 410 nm, 434nm, 486 nm and 656 nm) are shown below. Lines five and six can be seen with the naked eye but are considered to be UV (< 400 nm).

https://en.wikipedia.org/wiki/Balmer_series

HYDROGEN'S ATOMIC ENERGY LEVELS

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

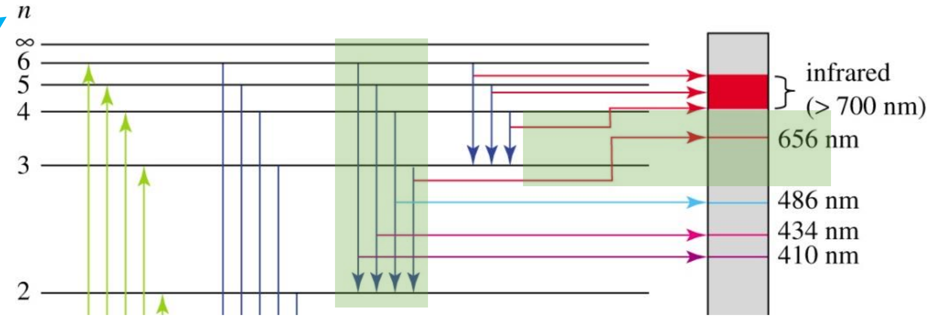


As energy levels approach $n = \infty$ the energy differences between them become smaller. Since **energy and wavelength are inversely proportional** the spectrum here has the lines converging *the opposite* to the energy level diagram.

For example the transition at 656 nm (red) is due to the smallest transition in the Balmer series

($E_n = 3 \rightarrow E_n = 2$).

Principal quantum number



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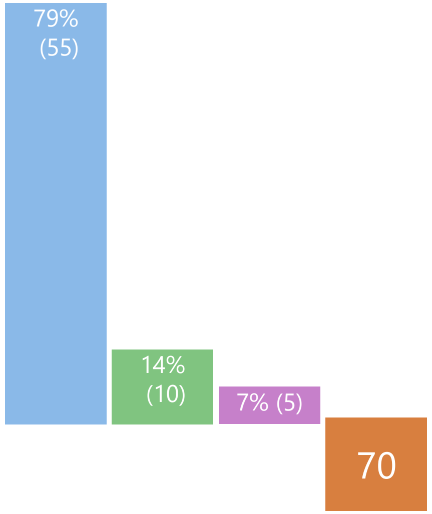
RECAP – ATOMIC SPECTRA

A single line on an emission spectra of the hydrogen atom represents __________ with energy equal to the difference between two energy levels within the atom being emitted from the atom.

- A — A photon
- B — An electron
- C — Either a photon or an electron



RYDBERG EQUATION

The Swedish physicist **Johannes Rydberg** (1854–1919) subsequently restated and expanded Balmer's result in the Rydberg equation:

Rydberg constant = 1.097 x 10^7 m^{-1}

$$\boxed{\bar{\nu}} = \frac{1}{\lambda} = -\boxed{R_H}\left(\frac{1}{{n_f}^2} - \frac{1}{{n_i}^2}\right)$$

wavenumber

$\frac{1}{\lambda}$ is another way of expressing wavenumber

By Per Bagge (1866-1936) - Original photograph in the collections of The Archives and Museum of the Academic Society in Lund, Sweden., Public Domain, https://commons.wikimedia.org/w/index.php?curid=41525001

This equation relates the wavelength of a photon absorbed or emitted to the transition between two energy levels of a hydrogen atom.

RYDBERG WORKED EXAMPLE

What is the wavelength of a photon absorbed to facilitate a n=1 to n=5 transition?

$$\begin{aligned} 1/\lambda &= -R_H \times (1/5^2 - 1/1^2) \\ &= -1.10 \times 10^7 \times (1/25 - 1) \\ &= 1.05 \times 10^7\ \mathrm{m}^{-1} \end{aligned}$$

$$\therefore \lambda = 9.49 \times 10^{-8}\ \mathrm{m}$$

$$(= 95\ \mathrm{nm})$$

IONISATION ENERGY

How much energy is required to completely remove an electron from a ground state hydrogen atom?

The removal of an electron from a neutral atom is called the ionisation energy.

For one H atom, $\Delta E = E_{final} - E_{initial}$

$$\Delta E = -\overset{*}{\left(2.18\times10^{-18}\text{ J}\right)}\times\left(\frac{1}{n_{final}^{2}}-\frac{1}{n_{initial}^{2}}\right) = -\left(2.18\times10^{-18}\text{ J}\right)\times(0-1)$$

$$\Delta E = 2.18\times10^{-18}\text{ J}$$

For a mole of H atoms, the molar ionisation energy would be :

$$\Delta E = \left(2.18\times10^{-18}\text{ J}\right)\times\left(6.023\times10^{23}\text{ mole}^{-1}\right) = 1.31\times10^{6}\text{ Jmol}^{-1}$$

* Rydberg constant in J instead of m^{-1}

RECAP

Light is

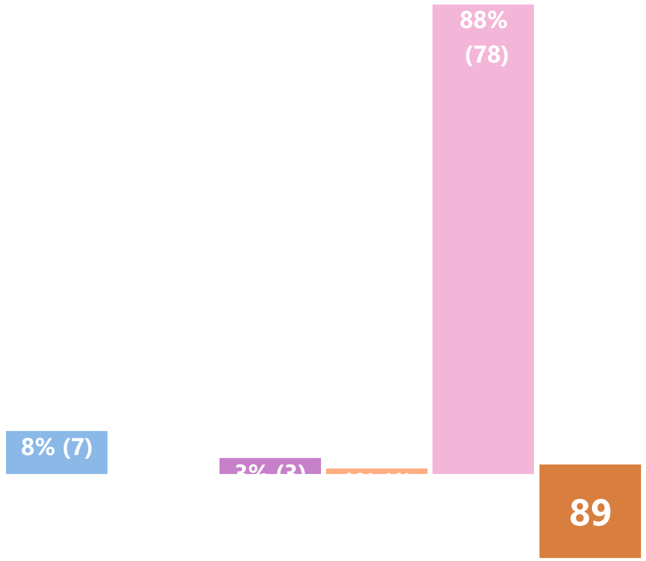
A — A wave

B — A particle

C — A photon

D — Electric and magnetic fields

E — All of the above

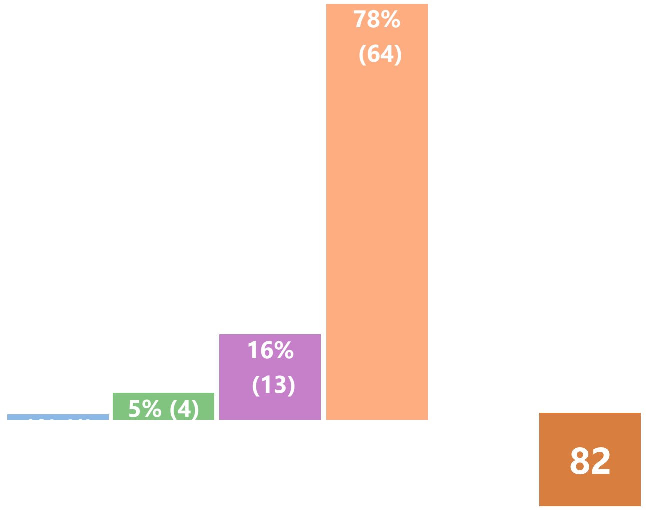


RECAP

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

The energy of a photon is proportional to its ____________ and inversely proportional to its ____________

- **A** Wavelength, velocity
- **B** Frequency, velocity
- **C** Wavelength, frequency
- ✓ **D** Frequency, wavelength
- **E** Don't know



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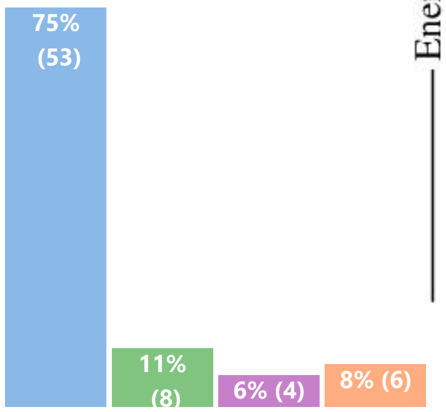
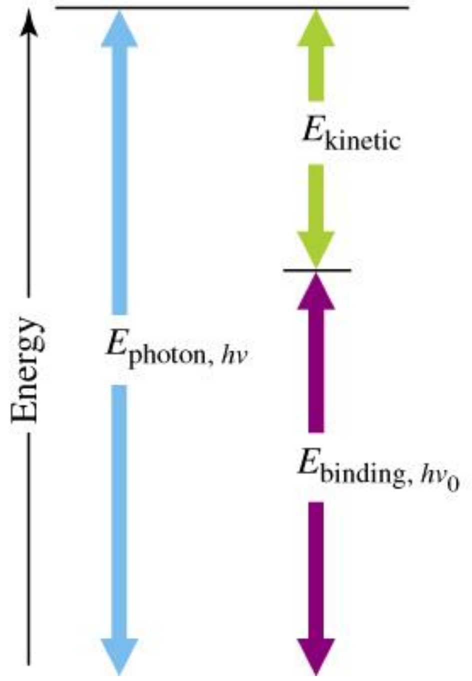
RECAP

If a photon has surplus energy to cause an electron to be ejected from a metal surface, what is the kinetic energy of the electron equal to?

- A $E_{photon} - E_{binding}$
- B $E_{binding} - E_{photon}$
- C E_{photon}
- D $E_{binding}$
- E Don't know





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RECAP

An electron can have a ____________ of energy *outside* an atom (kinetic) but only ____________ energy *inside* an atom.

- A Surplus, potential
- B Continuum, potential
- C Continuum, discrete
- D Discrete, continuum
- E Don't know

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