$= 2.65 \times 10^{-10}$ m $=$ $\underline{\underline{0.265 \text{nm}}}$

e.g. In a typical crystal, this distance is 0.25 nm. Determine the energy of a photon particle-wave beam with this wavelength and of an electron particle-wave beam with this wavelength.

photon energy $=$ $E = h\nu = hc/\lambda$

$$E_{photon} = \frac{(6.626 \times 10^{-34} Js)(2.998 \times 10^{8} m/s)}{(0.25 nm)(10^{-9} m/nm)}$$

$$= 7.9 \times 10^{-16} J$$

$$\lambda_{particle} = \frac{h}{mv}$$

$$E_{kinetic} = \frac{1}{2}mv^2$$

Find speed of electron:

$$V_{electron} = \frac{h}{m\lambda} = \frac{(6.626 \times 10^{-34} kg\, m^2/s)}{(9.109 \times 10^{-31} kg)(0.25 \times 10^{-9} m)}$$

$$= 2.91 \times 10^{6} m/s$$

Use speed to find the kinetic energy of the electron:

$$E_{kinetic} = \frac{1}{2}mv^2 = \frac{(9.109 \times 10^{-31} kg)(2.91 \times 10^{6} m/s)^2}{2}$$

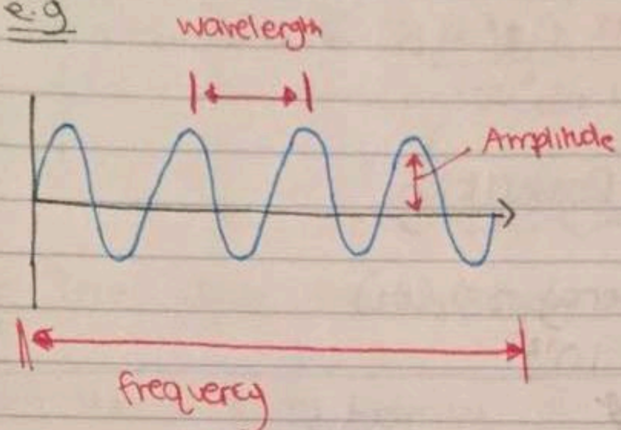
$$= 3.9 \times 10^{-18} kg\, m^{-2}$$

$$\boxed{= 3.9 \times 10^{-18} J}$$

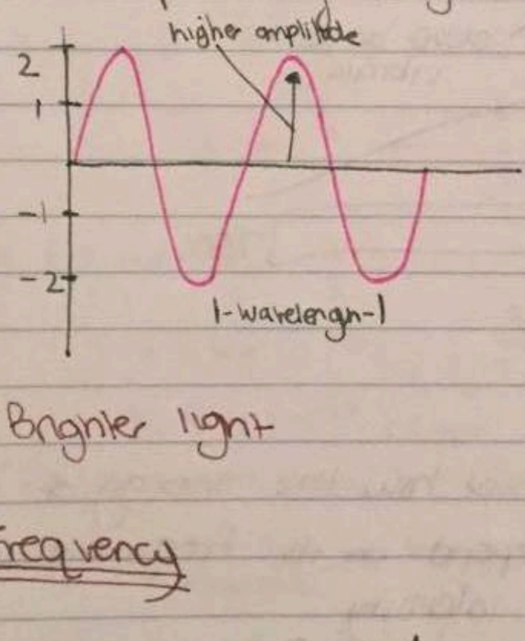
- amplitude - the height of a wave. ~~~~ The amplitude of a light wave measures the intensity of the light.

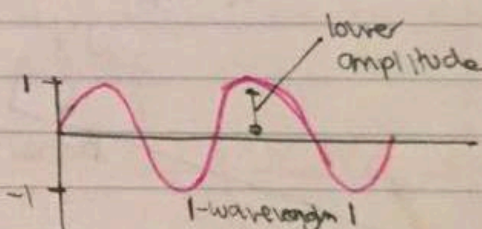
Speed of light: $c = 2.998 \times 10^{8}\ m/s$ ★ $3.00 \times 10^{8}\ ms^{-1}$

e.g



The amplitude of a light:





& Dim light

* Brighter light

Frequency

$$\nu = c / \lambda$$

frequency — wavelength — speed of light

* The lower the energy is, the more energy must be supplied to remove its electron.
* These are ~~any~~ energy changes that are measured relative to the energy of a free electron.
* The kinetic energy of a freely moving electron is positive relative to this conventional zero point.
* In contrast, bound electrons are lower in energy than free stationary electrons, so they have negative energy values relative to the zero point.
* ∴ "Most" stable = lowest or most negative
* a bound electron has "negative energy" only in relation to a free electron, because energy is released when a free stationary electron binds to an atom.

NOTE: Excited states are not stable

$$\boxed{\Delta E_{atom} = \pm h\nu_{photon}}$$

- when an atom absorbs a photon, atom gains photon's energy, ΔE_{atom} is positive.
- when an atom emits a photon, atom loses photons energy, ΔE_{atom} is negative.
- As an atom returns from an excited state to the ground state, it must lose exactly the amount of energy that it originally gained.

e.g. A sodium-vapor street emits yellow light at wavelength $\lambda = 589$ nm. What is the energy change for a sodium atom involved in this emission? How much energy is released per mole of sodium atoms?

→

(4) Each metal has its own characteristic threshold frequency because electrons are bound more tightly to some metals than to others.

eg → The Photoelectric effect!!

The minimum energy needed to remove an electron from a potassium metal surf is 3.7×10^{-19} J. Will photons of frequencies $4.3 \times 10^{14} s^{-1}$ (red light) and $7.5 \times 10^{14} s^{-1}$ (blue light) trigger the photoelectric effect? If so, what is the maximum kinetic energy of the ejected electrons?

$= E_{red\ photon} = h\nu = (6.626 \times 10^{-34} J\,s)(4.3 \times 10^{14} s^{-1}) = 2.8 \times 10^{-19} J$

$= E_{blue\ photon} = h\nu = (6.626 \times 10^{-34} J\,s)(7.5 \times 10^{14} s^{-1}) = 5.0 \times 10^{-19} J$

$$E_{kinetic\ (electron)} = h\nu - h\nu_0$$

$$E_{kinetic\ (electron)} = (5.0 \times 10^{-19} J) - (3.7 \times 10^{-19} J) = 1.3 \times 10^{-19} J$$

∴ red light has a lower freq than blue light.

4.3 - Absorption and Emmission Spectra

* When an atom absorbs a photon of sufficiently high energy, an electron is ejected, a process called photoionization

* When atoms gain energy but do not ionize, the atom is transformed to a higher energy state called an **excited state**.

* The lowest energy state of an atom, which is its most STABLE state, is called the **ground state**.

• Divide by Avogadro's number to find volume per atom.

$$V_{atom} = \frac{V_{molar}}{N_A} = \frac{13.00\ cm^3/mol}{6.022 \times 10^{23}\ atoms/mol}$$

$$= \underline{2.16 \times 10^{-23}}\ cm^3/atom$$

Isotopes

mass number (m) → $^{1}_{1}H$ ← atomic number (z)

• 1 proton, 1 electron = ("protium")

mass number (m) → $^{2}_{1}H$ ← atomic number (z)

• 1 proton, 1 neutron, 1 electron
• ("deuterium")

* Isotopes of an element have IDENTICAL chemical properties.

4.2 - Characteristics of Light

• Light has wave-like properties
• Wave - regular oscillation in some particular property
• Frequency (ν) - the number of wave crests passing a point in space in one second. (s^{-1} or Hertz (Hz)).
• Wavelength (λ) - Distance b/w successive wave crests. (m or nm)

4.7 - Sunlight and the Earth

Q) Assume gasoline is octane ($C_8H_{18}(l)$)
Find ΔH combustion:

$$C_8H_{18}(l) + 12.5O_2(g) \longrightarrow 8CO_2(g) + 9H_2O(g)$$

$$\Delta H_{combustion} = 8(-393) + 9(-285.6) - (-208.2)$$
$$= -5506.2 \text{ kJ/mol}$$

Other Energy Source

7. Coal Conversion:

- convert to gaseous fuels
- large molecules → small molecules
- must break C-C bonds
- C-H and C-O bonds are made

Coal Gasification:

Coal + Water + Air

↓

Methane + Synthesis Gas

Making syngas:

$$C(s) + H_2O(g) \rightarrow \underbrace{CO(g) + H_2(g)}_{\text{synthesis gas (syngas)}} \quad \Delta H = +131 \text{ kJ}$$

endothermic!

- syngas is a fuel on its own...

$\frac{1}{2} CO_{(g)} + \frac{1}{4} O_{2(g)} \rightarrow \frac{1}{2} CO_{2(g)}$ $\Delta H = -142 kJ$

$\frac{1}{2} H_{2(g)} + \frac{1}{4} O_{2(g)} \rightarrow \frac{1}{2} H_2O_{(g)}$ $\Delta H = -121 kJ$

syngas $+ \frac{1}{2} O_{2(g)} \rightarrow \frac{1}{2} CO_{2(g)} + \frac{1}{2} H_2O_{(g)}$

$\Delta H = -263 kJ$

... has a low fuel value

Upgrading syngas:

1. $CO_{(g)} + H_2O_{(g)} \rightarrow CO_{2(g)} + H_{2(g)}$
   ("CO shift reaction") $\Delta H = -41 kJ$
2. $CO_{2(g)}$ is removed
3. $CO_{(g)} + 3H_{2(g)} \rightarrow CH_{4(g)} + H_2O_{(g)}$
   $\Delta H = -206 kJ$
4. H_2O is removed, leaving Synthetic Natural Gas, SNG ($CH_{4(g)}$)

USES OF SYNGAS:

* Production of methanol
* Production of Hydrogen

2. Hydrogen:

$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(g)}$ $\Delta H = -242 kJ$

to produce hydrogen:

$H_2O_{(l)} \rightarrow H_{2(g)} + \frac{1}{2} O_{2(g)}$ $\Delta H = +286 kJ$

* Hydrogen is compressed and stored in a pressure tank
* Stored in a solid compound
* Cooled to a liquid state and kept cold in a properly insulated tank.

• When a metal surface absorbs a photon, the energy of the photon is transferred to an electron:

$$\Delta E_{electron} = E_{photon}$$

• Some of this energy is used to overcome forces that bind the electron to the metal, and the remainder shows up as kinetic energy of the ejected electron.

• The energy of a photon at the threshhold frequency equals to the minimum energy needed to overcome the forces that bind the electron to the metal.

$$\therefore \quad E_{kinetic\ (electron)} = hv - hv_0$$

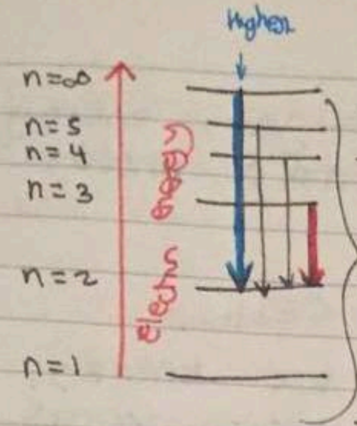
The observed properties of the photoelectric effect:

① When the energy of the photon is less than hv_0 (low-frequency light), there is not enough energy per photon to overcome the binding energy of the electron. Under these conditions, no electron can escape from the metal surface, no matter how intense the light.

② After the energy of the photon exceeds the threshold value ($hv > hv_0$), the electrons are ejected. The extra energy of the photon is transferred to the ejected electron as kinetic energy; this extra kinetic energy increase nearly within v.

③ The intensity of a light beam is a measure of the number of photons: light with higher amplitude comes more photons than light of lower amplitude. The intensity of light does not determine the amount of energy per photon.

– "Higher intensity" means <u>more</u> photons but <u>not</u> more energy per photon.



→ Higher energy goes from bottom to top
→ An electron with a LOW energy (e.g n=1) is close to the nucleus of the atom.
→ Balmer Series (VISIBLE) → saw a line that corresponded to the energy diff b/w the two levels.
→ Lyman series (UV) - realized that there are a bunch of lines that correspond to photons in UV portion of spectrum
→ All Lyman lines have higher energy than the Balmer lines, because going down to lower levels
→ Paschen Series (IR) - found a lower energy set of lines.

Balmer series (VIS): lowest energy level is $n=3 \rightarrow n=2$,

Lyman series (UV): is $n=2 \rightarrow n=1$

* Balmer - Rydberg Equation: Relates λ to upper and lower transition levels

$$\boxed{\frac{1}{\lambda} = R\left[\frac{1}{m^2} - \frac{1}{n^2}\right]}$$

$R = 0.01097\,nm^{-1}$

lower level

upper level

→ When an electron ~~energy~~ changes energy levels, the change in an electronic transition b/w quantum levels.
→ The change in energy of the atom is the diff b/w the two levels:

$$\Delta E_{atom} = E_{final} - E_{initial}$$

→ Photons always have positive energies, but energy changes (ΔE) can be positive or negative

- This energy change is negative because the atom loses energy. this lost energy appears as a photon whose energy is given by:

$$E_{photon} = |\Delta E_{atom}| = |-4.09 \times 10^{-19} J| = 4.09 \times 10^{-19} J$$

- To determine the wavelength of a photon:

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

$$\lambda_{photon} = \frac{hc}{E_{photon}} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ m/s)(10^{9}\ nm/m)}{(4.09 \times 10^{-19} J)}$$

$$\boxed{= 486\ nm}$$

4.4- Properties of electrons

* de Broglie equation *

moving massive particles exhibit wave-like properties:

$$\lambda_{particle} = \frac{h}{\underbrace{m}_{mass}\underbrace{v}_{velocity}}$$

e.g. An electron is moving at 2.74×10^{6} m/s. Find its wavelength.

$$\lambda = h/(mv)$$

$$= \frac{6.63 \times 10^{-34}\ Js}{9.11 \times 10^{-31}\ kg \times 2.74 \times 10^{6}\ ms^{-1}}$$

* Absorption spectrum - The absorption spectrum is unique for each gas because different atoms absorb different energies of photons. The gas in the tube absorbs light at specific wavelengths, called lines.
  - the intensity of transmitted light is low at these particular wavelengths.

* Emission Spectrum - A plot of the intensity of light emitted as a function of frequency

↳ The Emission Spectrum of Hydrogen:
- shows several sharp emission lines of high intensity.
- lines are 'atoms' ; called 'discreet lines'
- The frequencies of these lines correspond to photons emitted by the hydrogen atoms as they return to their ground state.
- Each frequency absorbed or emitted by an atom corresponds to a particular energy change for the atom.
- Each element has a unique emission pattern that provides valuable clues about atomic structure.

Quantization of Energy

- Electron energies are QUANTIZED
- Quantization - the energy levels of the electron can only have particular values.
  - specific discreet values.

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

- Energy is conserved, so the energy of the emitted photon must exactly equal the energy lost by the atom.

$$\Delta E_{photon} = -\Delta_{photon} = -\frac{hc}{\lambda}$$

- The equality includes a negative sign because the atom loses energy. To calc the energy of a photon:

$h = 6.626 \times 10^{-34}\ Js$, $c = 2.998 \times 10^{8}\ m/s$, $\lambda = 589\ nm = 589 \times 10^{-9}\ m$

$$\Delta E_{atom} = -E_{photon} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ m/s)}{(589 \times 10^{-9}\ m)}$$

$$= -3.37 \times 10^{-19}\ J$$

$$E_{molar} = (\Delta E_{atom})(N_A) = (-3.37 \times 10^{-19}\ J)(6.022 \times 10^{23}/mol)$$

$$= -2.03 \times 10^{5}\ J/mol$$
$$= -203\ kJ/mol$$

Atomic Spectra

- White light (composed of all colors) will pass through a slit (makes a straight line), then through a prism which disperses colors of the rainbow.
- Replacing white light with an Atomic Light source:
  - → consists of only one type of atom
  - → Starts with a tube with a very high voltage. Electrical energy travels through this and excites atoms. (electrons pushed to higher energy level)
  - → when electrons fall back down, their energy is converted to light.
  - → Through ~~the~~ slit, prism, then atomic light source give <u>DISCRETE</u> colors.

→ When absorption occurs, an atom gains energy, ΔE for the atom is positive, and a photon disappears:

$$E_{absorbed\ photon} = \Delta E_{atom}$$

→ When emission occurs, an atom loses energy, ΔE for the atom is negative and a photon appears:

$$E_{emitted\ photon} = -\Delta E_{atom}$$

* We can combine these two equations by using absolute values:

$$E_{photon} = |\Delta E_{atom}|$$

e.g. Hydrogen Energy Levels

Q. What is the energy change when the electron in a hydrogen atom changes from the fourth energy state to the second energy state? What is the wavelength of the photon emitted?

$$\Delta E_{atom} = E_{final} - E_{initial} = E_2 - E_4$$

$$E_n = -\frac{2.18 \times 10^{-18} J}{n^2}$$

$$E_2 = -\frac{2.18 \times 10^{-18} J}{2^2} = -5.45 \times 10^{-19} J$$

• The lower energy level has more negative energy. The energy diff is:

$$\Delta E_{atom} = E_{final} - E_{initial} = (-5.45 \times 10^{-19} J) - (-1.36 \times 10^{-19} J)$$
$$= -4.09 \times 10^{-19} J$$

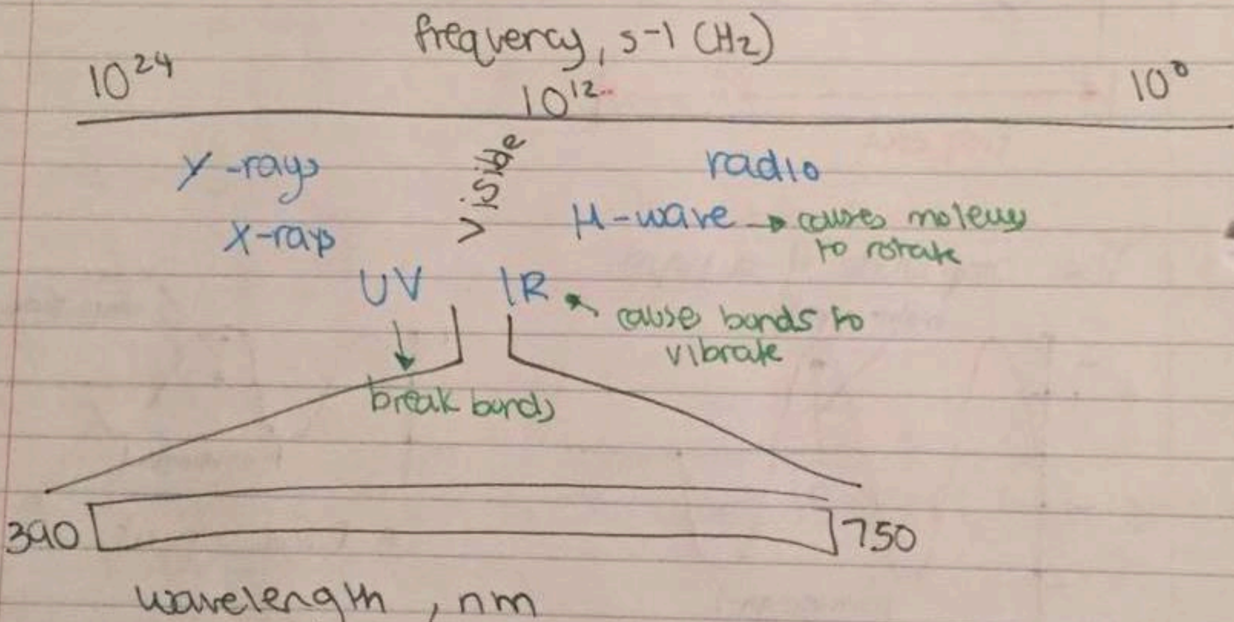
e.g. An FM radio station transmits its signal at 88.1 MHz. What is the wavelength of the radio signal?

$c = 2.998 \times 10^8 m/s \quad \nu = 88.1 MHz$

$\nu\lambda = c \quad so \ \lambda = \frac{c}{\nu}$

$$\lambda = \frac{2.998 \times 10^8 m/s}{88.1 \times 10^6 s^{-1}} = 3.40 m$$

Electromagnetic Radiation



* Photoelectric effect = shows how the energy of light depends on its frequency and intensity
 = The basis for many light-sensing devices, such as automatic door openers and camera exposure meters

different atoms have different binding energies

Heisenberg's Uncertainty Principle

- If the position (x) of an electron is known, its momentum (mv) is uncertain
- If its momentum is known, its position is uncertain:

$$(\Delta x)(\Delta mv) > \frac{h}{4\pi}$$

4.5 - Quantization & Quantum Numbers

Enter Erwin Schrodinger

"Atoms themselves are wavelike"

Schrodinger wave equation:

Don't need to know

The wave equation yields four quantum numbers:

magnetic

n, l, m_l, m_s

principal azimuthal spin

Principle Quantum Number:

n → size and energy level of orbital

$n = 1, 2, 3, 4, 5, 6 \ldots$

Azimuthal Quantum Number:

l → describes shape of orbital

$l = 0, 1, 2, \ldots, (n-1)$

(subshell number)

Subshells:

subshell no.	0	1	2	3
notation	s	p	d	f

* Light is also particle-like
* Light comes in packets or bundle, called photons.
* Each photon has an energy that is directly proportional to the frequency:

According to Planck:

photon enenergy (J)

$$E = h\nu$$

frequency (s^{-1})

planck's constant (6.63×10^{-34} Js)

* The larger the amplitude, the brighter the light.

e.g. What is the energy of a photon of red light of wavelength 655 nm?

$h = 6.626 \times 10^{-34}$ Js, $\lambda = 655$ nm, $c = 2.998 \times 10^{8}$

$E = h\nu$ and $\nu = \frac{c}{\lambda}$ so $E = \frac{hc}{\lambda}$

$$E_{photon} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ m/s)}{(655\ nm)(10^{-9}\ m/nm)} = 3.03 \times 10^{-19} J$$

eg $E = h\nu = h\left(\frac{c}{\lambda}\right)$

$= 6.63 \times 10^{-34}\ Js \left(\frac{3.00 \times 10^{8}\ ms^{-1}}{589 \times 10^{-9}\ m}\right)$

$= 3.4 \times 10^{-19}$ J (per photon)

$\times\ 6.02 \times 10^{23}\ mol^{-1} = 205{,}000$ J (mol photons)$^{-1}$

$= \underline{205\ kJ}$ (mol photons)$^{-1}$

∴ Light has wavelength properties and photon-light properties