Atomic Theory.

Finding The Energy of a Photon

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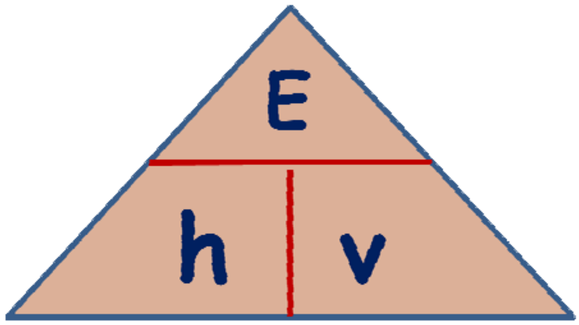
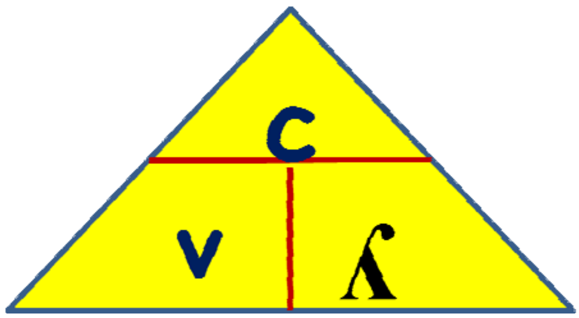
Some useful rearrangement triangles

The Relationship between light and energy



Converting frequency to wavelength



Also note that 1 mole = 6.02×10^{23}

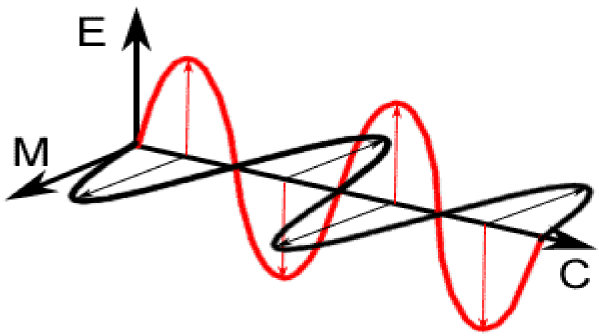
Planck's Equation $E = hv$

E stands for energy (in Joules),
v stands for frequency [in reciprocal seconds – written s^{-1} or Hertz (Hz)- 1Hz = 1 s^{-1}),
h is Planck's constant. Provided on your exam data sheet (**6.626×10^{-34} J.s**)

This equation is said to define the relationship between energy and frequency in a black body. A black body is an object that is both a perfect absorber and emmitter – absorbing all incident radiation and emitting all possible radiations. Black bodies do not need to be black. The Sun is an example of a black body.

The radical implication of Planck's equation was that light came in finite packets - multiples of hv.

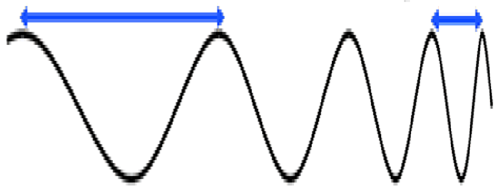
For the convenience of your future study of electromagnetic radiation, you might want to know the units often used for it.



For **FREQUENCY**>
1 Hz = 1 hertz = 1 cycle per second

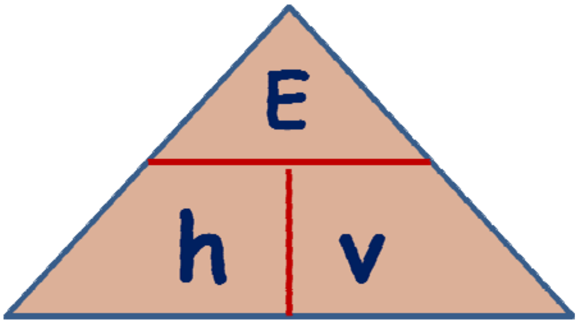
For **WAVELENGTH**
1 nm = 1 nanometer = 10^{-9} m:

We'll be considering wavelengths of IR, visible, UV and X-rays.



The highest energy waves have extemely short wavelength, where we use the pm unit

1 pm = 1 picometer = 10^{-12}m :



Using this rearrangement triangle, complete the following: -

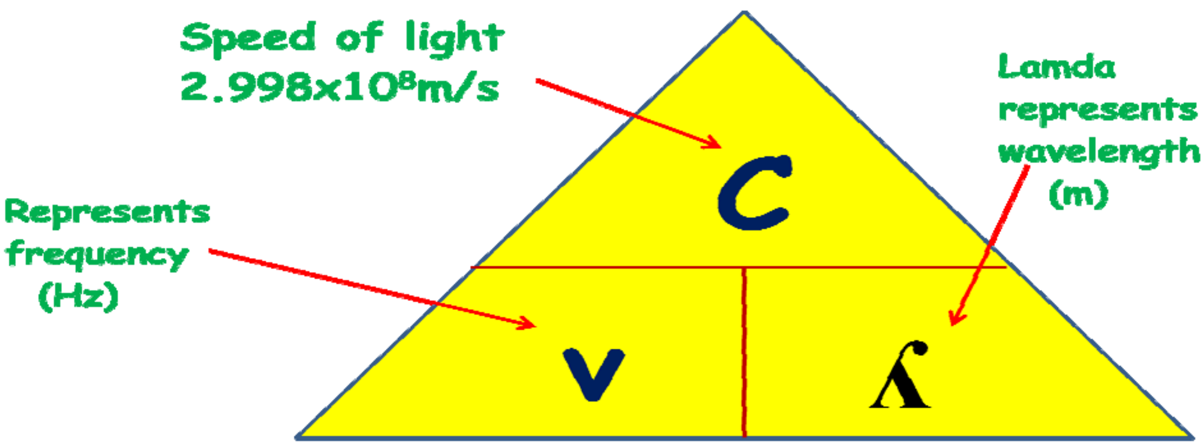
E =

h =

v =

Typically one is given the wavelength of the radiation. It is then necessary to use the other equation to convert the wavelength (ʎ) to frequency (v). N.B. **Wavelength in m**

Converting between frequency and wavelength.



v = frequency; C = speed of light; ʎ = wavelength
Hz (waves/s) 2.998x10^{8}m/s metres (m)

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It is possible to combine the two equations above. We see that $v = c/\lambda$

So in the equation $E = hv$ we can replace the v with c/λ.

This gives us $E = hc/\lambda$.

However, using each equation separately works fine. The typical question requires: -

Wavelength (nm) → Wavelength (m) → Frequency (Hz) → Energy of Photon (J) → Energy of 1 mole of photons (kJ)

Try the following questions.

1. Orange light has a wavelength of 620 nm.

(a) What is its wavelength in metres?

(b) What is its frequency?

(c) **Now use Plank's Law to work out the energy of one quantum of orange light.**

The solution is on page 5 and fully discussed via video at justchemy.com

2. Calculate the energies of one photon of light as follows: -

(a) ultraviolet ($\lambda = 2 \times 10^{-8}$ m),

(b) visible ($\lambda = 6 \times 10^{-7}$ m)

(c) infrared ($\lambda = 4 \times 10^{-4}$ m)

The solution is on page 5 and fully discussed via video at justchemy.com

3. Green light has a wavelength of approximately <u>500nm</u>. What is the frequency of this light. What is the energy in Joules of one photon of green light? What is the energy in Joules of 1.00 mole of photons of green light?

Recall, in very approximate terms, the wavelengths of coloured light.

Red	Orange	Yellow	Green	Blue	Indigo	Violet
750 nm			Incresing energy →			380 nm

The solution is on page 6 and fully discussed via video at justchemy.com

4. The yellow light emitted by the sodium lamps in streetlights has wavelengths of 589.6 nm and 589.0 nm. What is the frequency of the <u>589.0 nm light</u>? What is the energy of 1 mole of photons of yellow light with a wavelength of 589.0 nm?

The solution is on page 6 and fully discussed via video at justchemy.com

5. Find the frequency of a photon whose energy would be sufficient to

<u>cleave a Cl-Cl molecule.</u> Use the Planck equation and the following data: -

Planck's constant = 6.63×10^{-34} Js.

The Avogadro Constant = 6.02×10^{23} mol^{-1}

The Cl-Cl bond strength is 242 $kJmol^{-1}$

The solution is on page 6 and fully discussed via video at justchemy.com

Solutions:

1. (a) 1nm = 10^{-9}m so 6s20nm = 620 X 10^{-9}m = 6.2 X10^{-7}m

(b) $v = c / \lambda = 2.998 \times 10^{8}(m/s) / 6.2 \times 10^{-7}(m) = 4.8 \times 10^{14}\ s^{-1}$

(c) $E = h \times v$ **= $(6.626 \times 10^{-34}\ J.s) \times (4.84 \times 10^{14}\ s^{-1}) = 3.21 \times 10^{-19}$ J**

Multiply numbers and indices separately and then combine them.

[Note: **$6.626 \times 4.84 = 32.1$ <u>and</u> $10^{-34} \times 10^{14} = 10^{(-34+14)} = 10^{-20}$**]

2 (a) We can use the combined equation **$E = h.c/\lambda$**

$$\frac{hc}{\lambda} = \frac{(\qquad\qquad Js) \times (\qquad\qquad m/s)}{m}$$

where h = Planck's Constant = 6.626×10^{-34} Js,

c = speed of light in a vacuum = 3.00×10^{8} m/s, and λ = wavelength.

$$\frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ Js) \times (3.00 \times 10^{8}\ m/s)}{2 \times 10^{-8} m}$$

So for this wavelength of <u>ultraviolet light</u>, E = 9.939×10^{-18} J

Treating parts (b) and (c) in the same manner should yield: -

(b) For this wavelength of Visible Light E = 3.313×10^{-19} J

(c) For this wavelength of Infra red Light E = 4.9695×10^{-22} J

3. $\lambda = 500$ nm → $\lambda = 5 \times 10^{-7}$ m → $v = 6.0 \times 10^{14}$ Hz→ E = <u>**3.976**</u> $\times 10^{-19}$ J/photon →

$E = 2.38 \times 10^{5}$ J/mol→ E = 238,000J/mol →

<u>E = 238kJ/mol for this wavelength of GREEN light</u>

4. $\lambda = 589$ nm → $\lambda = 5.89 \times 10^{-7}$ m → $v = 5.09 \times 10^{14}$ Hz→ E = <u>**3.37**</u> $\times 10^{-19}$ J/photon → $E = 2.02 \times 10^{5}$ J/mol→ E = 202,000J/mol →

<u>E = 202kJ/mol for this wavelength of Yellow light</u>

As expected, green light has more energy than yellow light.

Red	Orange	Yellow	Green	Blue	Indigo	Violet
750 nm			Incresing energy →			380 nm

5. 242 kJ to break 1 mole of Cl-Cl bonds

How can we work out how much energy is needed to to break just one Cl-Cl bond?

$= 242/6.02 \times 10^{23}$ kJ to break just one Cl-Cl bond

$= 4.02 \times 10^{-24}$ kJ to break just one Cl-Cl bond

$= 4.02 \times 10^{-21}$ J to break just one Cl-Cl bond

$$E = hv \quad \text{so } v = E/h$$

$$v = \frac{4.02 \times 10^{-21}\ \text{J}}{6.63 \times 10^{-34}\ \text{J.s}} = \underline{6.0633 \times 10^{12}\ \text{Hz}}$$