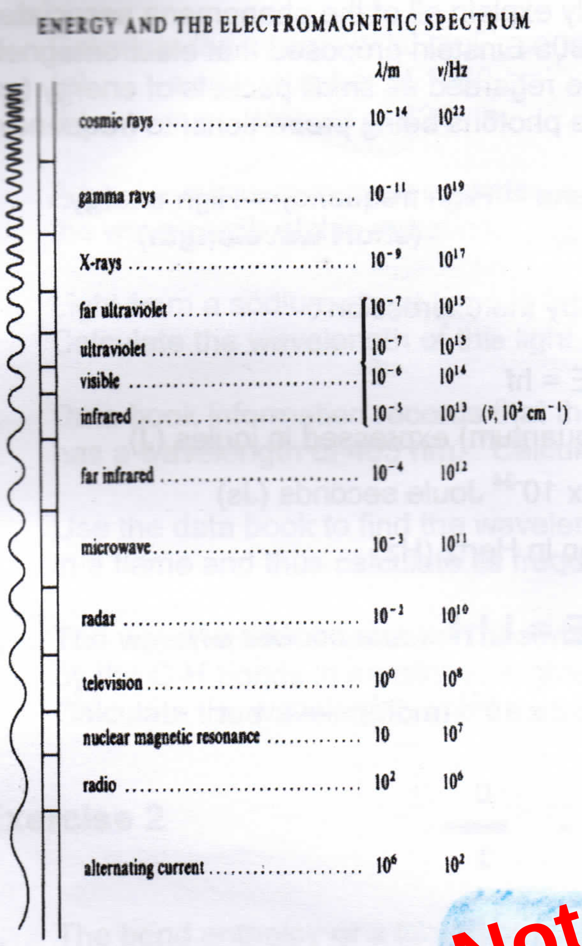

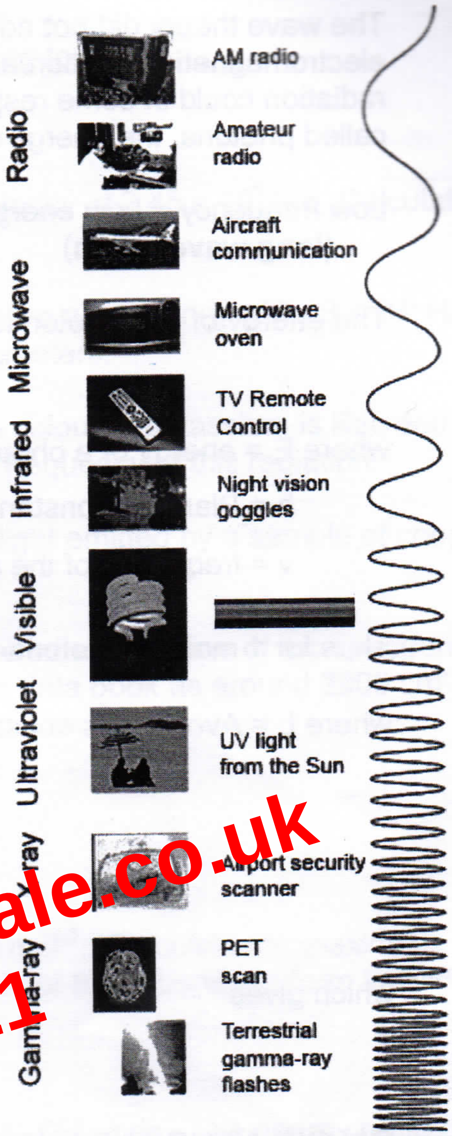

Electromagnetic spectrums can be presented in different ways, as seen above or on the previous page. No matter what way they are presented you will be expected to interpret them correctly.

QUANTUM THEORY OF RADIATION

The wave theory did not adequately explain all of the phenomena associated with electromagnetic radiation and in 1905 Einstein proposed that electromagnetic radiation could in some respects be regarded as small packets of energy (quanta) called photons, the energy of these photons being proportional to frequency.

Low frequency = Low energy (**long wavelength**) and High frequency = High energy. (**short wavelength**)

The energy of any photon is given by the expression :

$$E = hf$$

where E = energy of a photon (or quantum) expressed in joules (J)

h = Planck's constant, 6.63×10^{-34} Joule seconds (Js)

ν = frequency of the radiation in Hertz (Hz)

Thus for 1 mole of photons **E = Lhf** *(data booklet p4)*

where L = Avogadro's constant, 6.02×10^{23} (mol^{-1}) ---- but

$$\mathbf{f} = \frac{c}{\lambda}$$

which gives - $$\mathbf{E} = \frac{Lhc}{\lambda}$$ for 1 mole of photons of a given wavelength

CALCULATING THE ENERGY ASSOCIATED WITH ONE MOLE OF PHOTONS

For example, calculate the energy associated with one mole of photons of wavenumber 2000 cm^{-1}.

since $\bar{\nu} = \dfrac{1}{\lambda}$ wavenumber = $\dfrac{1}{\text{wavelength}}$

$\lambda = \dfrac{1}{2000} = 5 \times 10^{-4}$ cm or 5×10^{-6} m

Using the relationship -

$$\mathbf{E} = \frac{Lhc}{\lambda}$$

$$= \frac{6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 3 \times 10^{8} \times 10^{-3} \text{ (conversion factor for kJ)}}{5 \times 10^{-6}}$$

= **23.947 kJ.mol^{-1}**

The lines converge because the energy levels get closer together as the quantum numbers increase