(b) $\bar{\nu}$ (wavenumber) $= \frac{1}{\lambda} = \frac{1}{408 \times 10^{-9}\text{m}} = 2.45 \times 10^{6}\text{m}^{-1}$

(c) $\lambda = 408 \times 10^{-9}\text{m} \times \frac{10^{10}\text{Å}}{\text{m}} = 4080\text{Å}$

(d) $E = h\nu \times N_A = 6.63 \times 10^{-34} \times 7.353 \times 10^{14} \times 6.02 \times 10^{23}$ J/mole

$= 2.93 \times 10^{5}$ J/mole = 293 kJ/mole

(e) visible spectrum: violet (500 nm) red (800 nm) 408 nm = UV

**Problem #3**

For "yellow radiation" (frequency, $\nu$, $= 5.09 \times 10^{14}\ \text{s}^{-1}$) emitted by activated sodium, determine:

(a) the wavelength ($\lambda$) in [m]
(b) the wave number ($\bar{\nu}$) in [cm$^{-1}$]
(c) the total energy (in kJ) associated with 1 mole of photons

**Solution**

(a) The equation relating $\nu$ and $\lambda$ is $c = \nu\lambda$ where c is the speed of light $= 3.00 \times 10^{8}$ m.

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^{8}\ \text{m/s}}{5.09 \times 10^{14}\ \text{s}^{-1}} = 5.89 \times 10^{-7}\text{m}$$

(b) The wave number is 1/wavelength, but since the wavelength is in m, and the wave number should be in cm$^{-1}$, we first change the wavelength into cm:

$$\lambda = 5.89 \times 10^{-7}\ \text{m} \times 100\text{cm/m} = 5.89 \times 10^{-5}\ \text{cm}$$

Now we take the reciprocal of the wavelength to obtain the wave number:

$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{5.89 \times 10^{-5}\ \text{cm}} = 1.70 \times 10^{4}\ \text{cm}^{-1}$$

(c) The Einstein equation, $E = h\nu$, will give the energy associated with one photon since we know h, Planck's constant, and $\nu$. We need to multiply the energy obtained by Avogadro's number to get the energy per mole of photons.

$$h = 6.62 \times 10^{-34}\ \text{J.s}$$

$$\nu = 5.09 \times 10^{14}\ \text{s}^{-1}$$

$$E = h\nu = (6.62 \times 10^{-34}\ \text{J.s}) \times (5.09 \times 10^{14}\ \text{s}^{-1}) = 3.37 \times 10^{-19}\ \text{J per photon}$$