1) We are told that the threshold energy for titanium is $6.94 \times 10^{-19}$ J. We are also told that the wavelength of light used is $283$ nm ($= 283 \times 10^{-9}$ m), so we can get the energy of the light, and from that, the excess energy beyond the work function ($= \text{threshold energy}$), which is kinetic energy.

$$E = \frac{hc}{\lambda} = \frac{(6.6261 \times 10^{-34} \text{ J} \cdot \text{s})(2.9979 \times 10^8 \text{ m/s})}{283 \times 10^{-9} \text{ m}} = 7.02 \times 10^{-19} \text{ J},$$

which represents $E_k + E_p$, so $E_k = 7.92 \times 10^{-21} \text{ J} = \frac{1}{2} mv^2$. This leads to $v = \sqrt{\frac{2E_k}{m_e}} = \sqrt{\frac{1.58 \times 10^{-20} \text{ J}}{9.1094 \times 10^{-31} \text{ kg}}} = 1.32 \times 10^5 \text{ m/s}$.

2) (a real give-away) If $n=8$, then $l$, which can vary from 0 to $(n-1)$, has values of 0, 1, 2, 3, 4, 5, 6, 7 which would correspond to sub-shells of: one 8s, three 8p, five 8d, seven 8f, nine 8g, eleven 8h, thirteen 8i, and fifteen 8j. (Note that not all of these need to be occupied, and in fact, as we will see, many of them will not be, at least not in the foreseeable future!)

3) We are told that the energy required to dissociate the average C-C bond is 347 kJ/mol, which is equivalent to $347000 \text{ J/mol}$.

Using: $E = \frac{hc}{\lambda}$, which equation, upon rearrangement, yields $\lambda = \frac{hc}{E}$. Substituting and solving yields: $\lambda = \frac{(6.6262 \times 10^{-34} \text{ J} \cdot \text{s})(2.9979 \times 10^8 \text{ m/s})}{5.762 \times 10^{-19} \text{ J}} = 3.447 \times 10^{-7} \text{ m} = 344.7 \text{ nm}$. Since we are also asked for the frequency, this would be $v = \frac{2.9979 \times 10^8 \text{ m/s}}{3.447 \times 10^{-7} \text{ m}} = 8.696 \times 10^{14} \text{ s}^{-1}$. This wavelength and frequency put the light into the ultraviolet region of the electromagnetic spectrum.
4) We have a wavelength, \( \lambda \), of \( 1.55 \times 10^{-2} \) m. A photon of this light would have energy of: \[ E = \left( \frac{6.6261 \times 10^{-34} \text{ J} \cdot \text{s}}{2.9979 \times 10^8 \text{ m/s}} \right) \frac{1.55 \times 10^{-2} \text{ m}}{1} = 1.282 \times 10^{-23} \text{ J}. \]

We also know that the temperature of the water increases from 20.0°C to 98.0°C, or 78.0°C. Using \( q = C_p(\text{mass})(\Delta T) \), we get: \[ q = (252 \text{ g})(4.184 \text{ J/g} \cdot \text{°C})(78.0 \text{ °C}) = 82241 \text{ J}. \]

If each photon has energy \( 1.282 \times 10^{-23} \) J, and we need 82241 J to do our heating, we need: \[ \frac{82241 \text{ J}}{1.282 \times 10^{-23} \text{ J/photons}} = 6.415 \times 10^{27} \text{ photons}. \] (This is about 10650 moles of photons!)

5) If the emission was of yellow light, it seems fairly reasonable to assume that absorption of yellow light would also cause the excitation to the n=5 state. Yes, it is true that blue light would be of higher energy than the yellow light and could also cause an excitation; however, it would probably not be to the n=5 state. What is clear, though, is that red light, being of lesser energy than the yellow light, could not do the desired transition.