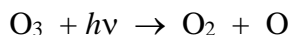


Light Energy, Bond Energy, and Gas Laws: An Integrative Problem

Problem 8-101

Ozone, O₃, absorbs ultraviolet radiation and dissociates into O₂ molecules and O atoms:



A 1.00-L sample of air at 22°C and 748 mmHg contains 0.25 ppm of O₃. How much energy, in joules, must be absorbed if all the O₃ molecules in the sample of air to dissociate? Assume that each photon absorbed causes one O₃ molecule to dissociate, and that the wavelength of the radiation is 254 nm.

Solution:

From chapter 6, we might be led to assume that the concentration of O₃ is in parts-per-million by volume [V(O₃)/10⁶V(air)]. So,

1. Calculate the quantity, in moles, of O₃ using the ideal gas law
2. Calculate the energy of 254 nm electromagnetic radiation in J/mol
3. Calculate the total energy required to dissociate the molar quantity of O₃

First, the mole amount of air:

$$n = \frac{PV}{RT} = \frac{\left(\frac{748 \text{ mmHg}}{760 \frac{\text{mmHg}}{\text{atm}}}\right)(1.00 \text{ L})}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(295.2 \text{ K})} = 0.04063 \text{ mol air}$$

Next, the mole amount of O₃ in the air. This is just a Dalton's Law or mole fraction problem:

$$n_{\text{O}_3} = 0.04063 \text{ mol air} \times \frac{0.25 \text{ mol O}_3}{10^6 \text{ mol air}} = 1.016 \times 10^{-8} \text{ mol O}_3$$

Calculate the energy of 254 nm in J/mol:

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \frac{\text{m}}{\text{s}})}{254 \times 10^{-9} \text{ m}} \times (6.022 \times 10^{23} \frac{\text{photons}}{\text{mol}}) = 471,300 \frac{\text{J}}{\text{mol}}$$

Finally, calculate the total energy required:

$$E = (471,300 \frac{\text{J}}{\text{mol}})(1.016 \times 10^{-8} \text{ mol O}_3) = \boxed{4.79 \times 10^{-3} \text{ J}}$$

A second, equally interesting question is: "From this data, what is the approximate bond energy for the O-O bond in ozone?"