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

Problem 8-101

Ozone, O_3, absorbs ultraviolet radiation and dissociates into O_2 molecules and O atoms:

$$O_3 + h\nu \rightarrow O_2 + O$$

A 1.00-L sample of air at 22°C and 748 mmHg contains 0.25 ppm of O_3. How much energy, in joules, must be absorbed if all the O_3 molecules in the sample of air to dissociate? Assume that each photon absorbed causes one O_3 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 O_3 is in parts-per-million by volume [$V(O_3)/10^6V(\text{air})$]. So,

1. Calculate the quantity, in moles, of O_3 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 O_3

First, the mole amount of air:

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

Next, the mole amount of O_3 in the air. This is just a Dalton's Law or mole fraction problem:

$$n_{O_3} = 0.04063\ \text{mol air} \times \frac{0.25\ \text{mol}\ O_3}{10^6\ \text{mol air}} = 1.016 \times 10^{-8}\ \text{mol}\ O_3$$

Calculate the energy of 254 nm in J/mol:

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

Finally, calculate the total energy required:

$$E = \left(471{,}300\ \frac{\text{J}}{\text{mol}}\right)\left(1.016 \times 10^{-8}\ \text{mol}\ O_3\right) = \boxed{4.79 \times 10^{-3}\ \text{J}}$$

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