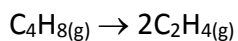


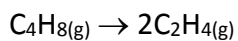
Answer on Question #73718 – Chemistry – General Chemistry

The following reaction has an activation energy of 262 kJ/mol



At 600.0 K the rate constant is $6.1 \times 10^{-8} \text{ s}^{-1}$. What is the value of the rate constant at 760.0 K.

Solution:



$$\log \frac{k_1}{k_2} = \frac{E_A}{2.303R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\log \frac{k_1}{k_2} = \frac{262000 \text{ J/mol}}{8.314 \text{ J} \cdot \text{mol/K} \times 2.303} \left(\frac{1}{760} - \frac{1}{600} \right)$$

$$\log \frac{k_1}{k_2} = -4.81$$

$$\frac{k_1}{k_2} = 10^{-4.81} = 1.55 \times 10^{-5}$$

$$k_2 = \frac{k_1}{1.55 \times 10^{-5}} = \frac{6.1 \times 10^{-8}}{1.55 \times 10^{-5}} = 3.9 \times 10^{-3} \text{ s}^{-1}$$

Answer provided by AssignmentExpert.com