Answer on Question #73718 – Chemistry – General Chemistry

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

$C_4H_{8(g)} \rightarrow 2C_2H_{4(g)}$

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

Solution:

$C_4H_{8(g)} \rightarrow 2C_2H_{4(g)}$

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

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

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

$$\frac{k_1}{k_2} = 10^{-4.81} = 1.55\times10^{-5}$$

$$k_2 = \frac{k_1}{1.55\times10^{-5}} = \frac{6.1\times10^{-8}}{1.55\times10^{-5}} = 3.9\times10^{-3}s^{-1}$$

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