Answer on the question #55936 - Chemistry - General chemistry

Question:

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

$C_4H_8(g)$ ---> $2C_2H_4(g)$

At 600.0 K the rate constant is 6.1*10^-8s^-1. What is the value of the rate constant at 750.0 K?

Solution:

According to Arrhenius' law, the rate constant is dependent on the activation energy and temperature as follows:

$$k = Ae^{-\frac{E_a}{RT}}$$

Then, the rate constants ratio at different temperature will be:

$$\frac{k_1}{k_2} = e^{-\frac{E_a}{R}(\frac{1}{T_1}-\frac{1}{T_2})}$$

Let's say $k_1$is the rate constant at 750 K and k_2 is the rate constant at 600 K. E_a is the activation energy, $6.1*10^{-8}$ s^{-1} hence:

$$k_1 = k_2e^{-\frac{E_a}{R}(\frac{1}{T_1}-\frac{1}{T_2})} = 6.1 \cdot 10^{-8} \cdot e^{-\frac{262 \cdot 10^3}{8.314} \cdot (\frac{1}{750}-\frac{1}{600})} = 2.2 \cdot\ 10^{-3}\ s^{-1}$$

Answer: $2.2*10^{-3}$ s^{-1}