The rate of a reaction doubles when its temperature changes from 300 K to 310 K. Activation energy of such a reaction will be (R = 8.314KJmol^{-1} and log 2=0.301)

(1) 53.6 KJmol^{-1}

(2) 48.6 KJmol^{-1}

(3) 58.5 KJmol^{-1}

(4) 60.5 KJmol^{-1}

Solution:

According to Arrhenius equation,

$\ln k_2/k_1 = -(E_a/2.303R)(1/T_2 – 1/T_1)$

$r_2/r_1 = k_2/k_1 = 2$

Given $T_1 = 300K$

$T_2 = 310\ K$

$\log(2) = -(E_a/2.303 \times 8.314)\ [(1/310) – (1/300)]$

$0.301 = -E_a/19.147\ [(300-310)/93000\ ]$

$0.301 = -E_a/19.147\ [-10/93000\ ]$

$E_a = 0.301 \times 19.147 \times 9300$

= 53598.19 J

= 53.6 kJ/mol