QUESTION

The rate of a reaction doubles when its temperature changes from 300 K to 310 K. Activation energy of such a reaction will be:

$(R = 8.314 JK^{-1} mol^{-1}$ and $\log_{10}2 = 0.301)$

A $58.5 kJ mol^{-1}$

B $60.5 kJ mol^{-1}$

C $53.6 kJ mol^{-1}$

D $48.6 kJ mol^{-1}$

ANSWER :

Correct option is C)

$\because \frac{r_2}{r_1} = \frac{k_2}{k_1} = 2$

Now $2.303 \log_{10}\frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$

$2.303 \log_{10} 2 = \frac{E_a}{8.314} \times \left[\frac{310-300}{310 \times 300}\right]$

$\therefore E_a = \frac{2.303 \times 0.301 \times 8.314 \times 310 \times 300}{10} = 53598.6 J$

$= 53.6 kJ mol^{-1}$