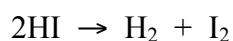


- At a certain temperature the following data were collected for the decomposition of HI.

Marks
4



Experiment	Initial [HI] (mol L ⁻¹)	Initial rate of reaction (mol L ⁻¹ s ⁻¹)
1	1.0×10^{-2}	4.0×10^{-6}
2	2.0×10^{-2}	1.6×10^{-5}
3	3.0×10^{-2}	3.6×10^{-5}

Determine the rate law for the reaction.

Between experiment (1) and (2), the concentration of HI is doubled. This leads to the rate increasing by a factor of 4.

Between experiment (1) and (3), the concentration of HI is trebled. This leads to the rate increasing by a factor 9.

The rate is proportional to [HI]²:

$$\text{rate} = k[\text{HI}]^2$$

What is the value of the rate constant for the decomposition of HI?

Using experiment (1), [HI] = 1.0×10^{-2} mol L⁻¹ and rate = 4.0×10^{-6} mol L⁻¹ s⁻¹:

$$\text{rate} = k[\text{HI}]^2$$

$$4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ s}^{-1} = (k) \times (1.0 \times 10^{-2} \text{ mol L}^{-1})^2$$

$$k = (4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ s}^{-1}) / (1.0 \times 10^{-2} \text{ mol L}^{-1})^2 = 4.0 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$$

Answer: $4.0 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$