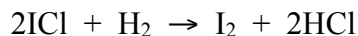


- At a certain temperature the following data were collected for the reaction shown.



Experiment	Initial [ICl] (mol L <sup>-1</sup> )	Initial [H <sub>2</sub> ] (mol L <sup>-1</sup> )	Rate of formation of [I <sub>2</sub> ] (mol L <sup>-1</sup> s <sup>-1</sup> )
1	0.10	0.10	0.0015
2	0.20	0.10	0.0030
3	0.10	0.050	0.00075

Determine the rate law for the reaction.

Between experiments (1) and (2), [ICl] is doubled and [H<sub>2</sub>] is constant. This change leads to a doubling of the rate: the rate is proportional to [ICl].

Between experiments (1) and (3), [ICl] is constant and [H<sub>2</sub>] is halved. This change leads to a halving of the rate: the rate is proportional to [H<sub>2</sub>].

Overall:

$$\text{rate} = k[\text{ICl}][\text{H}_2]$$

What is the value of the rate constant?

Using experiment (1), rate = 0.0015 mol L<sup>-1</sup> s<sup>-1</sup>, [ICl] = 0.10 mol L<sup>-1</sup> and [H<sub>2</sub>] = 0.10 mol L<sup>-1</sup>:

$$\text{rate} = k[\text{ICl}][\text{H}_2] = k(0.10 \text{ mol L}^{-1})(0.10 \text{ mol L}^{-1}) = 0.0015 \text{ mol L}^{-1} \text{ s}^{-1}$$

$$k = (0.0015 \text{ mol L}^{-1} \text{ s}^{-1}) / (0.10 \text{ mol L}^{-1})(0.10 \text{ mol L}^{-1}) = 0.15 \text{ L mol}^{-1} \text{ s}^{-1}$$

Answer: 0.15 L mol<sup>-1</sup> s<sup>-1</sup>

**THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.**

**Marks**  
**4**