Marks 4

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

$$2ICl + H_2 \rightarrow I_2 + 2HCl$$

Experiment	Initial [ICl] (mol L^{-1})	Initial [H_2] (mol L^{-1})	Rate of formation of [I_2] (mol L^{-1} s^{-1})
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_2] 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_2] is halved. This change leads to a halving of the rate: the rate is proportional to [H_2].

Overall:

rate = k[ICl][H_2]

What is the value of the rate constant?

Using experiment (1), rate = 0.0015 mol L^{-1} s^{-1}, [ICl] = 0.10 mol L^{-1} and [H_2] = 0.10 mol L^{-1}:

rate = k[ICl][H_2] = k(0.10 mol L^{-1})(0.10 mol L^{-1}) = 0.0015 mol L^{-1} s^{1}

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

Answer: **0.15 L mol^{-1} s^{-1}**

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