Marks 4

- The following data were obtained for the reaction of iodine atoms in the gas phase in the presence of argon.

$$2I(g) + Ar(g) \rightarrow I_2(g) + Ar(g)$$

Experiment Number	Initial [I] (M)	Initial [Ar] (M)	Initial Reaction Rate $-d[I(g)]/dt$ (M s^{-1})
1	1.0×10^{-5}	1.0×10^{-3}	8.70×10^{-4}
2	2.0×10^{-5}	1.0×10^{-3}	3.48×10^{-3}
3	2.0×10^{-5}	5.0×10^{-3}	1.74×10^{-2}

Derive an expression for the rate law for the formation of $I_2(g)$ and calculate the value of the rate constant for this reaction.

Between experiments (1) and (2), [Ar] is constant and [I] is doubled. This leads to the rate increasing by $\frac{3.48 \times 10^{-3}}{8.70 \times 10^{-4}}$: a factor of 4. The rate is proportional to $[I]^2$.

Between experiments (2) and (3), [I] is constant and [Ar] is increased by a factor of 5. This leads to the rate increasing by $\frac{1.74 \times 10^{-2}}{3.48 \times 10^{-3}}$: a factor of 5. The rate is proportional to $[Ar]^1$.

Overall:

$$-d[I(g)]/dt = k[I]^2[Ar]$$

From experiment (1), rate = 8.70×10^{-4} M s^{-1} when [I] = 1.0×10^{-5} M and [Ar] = 1.0×10^{-3} M. Hence:

$$8.70 \times 10^{-4} \text{ M s}^{-1} = k \times (1.0 \times 10^{-5} \text{ M})^2 \times (1.0 \times 10^{-3} \text{ M})$$

$$k = 8.70 \times 10^{9} \text{ M}^{-2} \text{ s}^{-1}$$

Rate law: **$-d[I(g)]/dt = k[I]^2[Ar]$**

Rate constant: **8.70×10^{9} M^{-2} s^{-1}**

Calculate the rate of appearance of $I_2(g)$ when $[I(g)] = 1.0 \times 10^{-3}$ M and $[Ar(g)] = 1.0 \times 10^{-2}$ M.

$$-d[I(g)]/dt = k[I]^2[Ar]$$
$$= (8.70 \times 10^{9} \text{ M}^{-2} \text{ s}^{-1}) \times (1.0 \times 10^{-3} \text{ M})^2 \times (1.0 \times 10^{-2} \text{ M}) = 87 \text{ M s}^{-1}$$

From the chemical equation, two I are lost for every I_2 produced. Hence:

$$d[I_2(g)]/dt = ½ \times -d[I(g)]/dt = ½ \times (87 \text{ M s}^{-1}) = 44 \text{ M s}^{-1}$$

Answer: **44 M s^{-1}**