Marks
2

- The following data were obtained for the reaction between gaseous nitric oxide and chlorine at –10 °C:

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

Experiment Number	Initial P_{NO} (atm)	Initial P_{Cl_2} (atm)	Initial Reaction Rate (atm s^{-1})
1	2.16	2.16	0.065
2	2.16	4.32	0.130
3	4.32	4.32	0.518

Derive an expression for the rate law for this reaction and calculate the value of the rate constant.

Between experiments (1) and (2), the partial pressure of NO is kept constant and the partial pressure of Cl_2 is doubled. This leads to a doubling of the rate so the reaction is first order with respect to Cl_2.

Between experiments (2) and (3), the partial pressure of Cl_2 is kept constant and the partial pressure of NO is doubled. This leads to the rate quadrupling. The reaction is second order with respect to NO.

Overall,

$$\textbf{rate} = kP_{NO}^2 P_{Cl_2} \quad \textbf{or rate} = k'[NO(g)]^2[Cl_2(g)]$$

As the partial pressure is proportional to the concentration. Either form is acceptable.

Using experiment (1), rate = 0.065 atm s^{-1} when P_{NO} = 2.16 atm and P_{Cl_2} = 2.16 atm. Hence,

$$0.065 \text{ atm s}^{-1} = k \times (2.16 \text{ atm})^2 \times (2.16 \text{ atm})$$

$$k = 0.0064 \text{ atm}^{-2} \text{ s}^{-1}$$

The units of *k* are obtained by ensuring that those in the equation balance:

$$\text{atm s}^{-1} = (\text{units of } k) \times (\text{atm})^2 \times (\text{atm}) \quad \text{so } k \text{ has units of atm}^{-2} \text{ s}^{-1}.$$

Rate law: **rate = $kP_{NO}^2 P_{Cl_2}$ or rate = $k'[NO(g)]^2[Cl_2(g)]$**

Rate constant: **k = 0.0064 atm^{-2} s^{-1}**