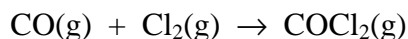


- Phosgene is a toxic gas prepared by the reaction of carbon monoxide with chlorine.



The following data were obtained in a kinetics study of its formation at 150 °C.

Experiment	Initial [CO] (mol L <sup>-1</sup> )	Initial [Cl <sub>2</sub> ] (mol L <sup>-1</sup> )	Initial rate (mol L <sup>-1</sup> s <sup>-1</sup> )
1	1.00	0.100	$1.29 \times 10^{-3}$
2	0.100	0.100	$1.33 \times 10^{-4}$
3	0.100	1.00	$1.30 \times 10^{-3}$
4	0.100	0.0100	$1.32 \times 10^{-5}$

Write the rate law for the formation of phosgene at 150 °C.

Between experiments (1) and (2), [Cl<sub>2</sub>] is constant and [CO] is increased by a factor of ten. This leads to a tenfold increase in the rate. The reaction is first order with respect to [CO].

Between experiments (2) and (3), [CO] is constant and [Cl<sub>2</sub>] is increased by a factor of ten. This leads to a tenfold increase in the rate. The reaction is also first order with respect to [Cl<sub>2</sub>].

Hence, the rate law is,

$$\text{rate} = k[\text{CO(g)}][\text{Cl}_2\text{(g)}]$$

Calculate the value of the rate constant at 150 °C.

In experiment (1), rate =  $1.29 \times 10^{-3}$  mol L<sup>-1</sup> s<sup>-1</sup> when [CO] = 1.00 mol L<sup>-1</sup> and [Cl<sub>2</sub>] = 0.100 mol L<sup>-1</sup>. Substituting into the rate law gives,

$$(1.29 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}) = k \times (1.00 \text{ mol L}^{-1}) \times (0.100 \text{ mol L}^{-1})$$

$$k = 1.29 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$$

The same method gives  $k = 1.33 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$ ,  $1.30 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$  and  $1.32 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$  for experiments (2), (3) and (4) respectively. The accuracy of the experiments suggests a value of  $1.3 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$ .

$$\text{Answer: } 1.3 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}$$

ANSWER CONTINUES ON THE NEXT PAGE

Calculate the rate of appearance of phosgene when  $[\text{CO}] = [\text{Cl}_2] = 1.3 \text{ M}$ .

Using the rate law derived above.

$$\begin{aligned}\text{rate} &= (1.3 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}) \times [\text{CO}] \times [\text{Cl}_2] \\ &= (1.3 \times 10^{-2} \text{ mol}^{-1} \text{ L s}^{-1}) \times (1.3 \text{ mol L}^{-1}) \times (1.3 \text{ mol L}^{-1}) \\ &= 2.2 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}\end{aligned}$$

Answer:  $2.2 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$