Marks 5

- A mixture of 0.500 mol of $NO_2(g)$ and 0.500 mol of $N_2O_4(g)$ is allowed to reach equilibrium in a 10.0 L vessel maintained at 298 K. The equilibrium is described by the equation below. $\Delta H° = -15$ kJ mol^{-1} for the forward reaction.

$$2NO_2(g) \rightleftharpoons N_2O_4(g) \qquad K_c = 1.2 \times 10^2\ M^{-1}$$

Show that the system is at equilibrium when the concentration of $NO_2(g)$ is 0.023 M.

The concentrations of $NO_2(g)$ and $N_2O_4(g)$ at the start are:

$$[NO_2(g)] = [N_2O_4(g)] = \frac{\text{number of moles}}{\text{volume}} = \frac{0.500\,\text{mol}}{10.0\,\text{L}} = 0.0500\,\text{M}$$

$[NO_2(g)]$ decreases during the reaction and so $[N_2O_4(g)]$ increases. From the chemical equation, *one* mole of $N_2O_4(g)$ is produced for every *two* moles of NO(g) that are lost.

The change in $[NO_2(g)]$ = (0.0500 – 0.023) M = 0.027 M. Hence,

$$[N_2O_4(g)]_{\text{equilibrium}} = (0.0500 + \tfrac{1}{2} \times 0.027)\ \text{M} = 0.064\ \text{M}$$

With these concentrations, the reaction quotient, *Q*, is given by:

$$Q = \frac{[N_2O_4(g)]}{[NO_2(g)]^2} = \frac{(0.064)}{(0.023)^2} = 120 = 1.2 \times 10^2$$

As *Q* = *K*, the reaction is at equilibrium.

Discuss the effect an increase in temperature, at constant volume, would have on the concentration of $NO_2(g)$.

As $\Delta H° = -15$ kJ mol^{-1} for the forward reaction, the reaction is exothermic. If the temperature is increased, the system will respond by removing heat. It will do this by shifting towards the reactant ($NO_2(g)$) as the backward reaction is endothermic. Hence, $[NO_2(g)]$ will increase.

State with a brief reason whether the concentration of $NO_2(g)$ is increased, decreased, or unchanged when argon gas (0.2 mol) is injected while the temperature and volume remain constant.

Adding argon will increase the pressure inside the vessel will increase. However, the inert gas does not change the volume so all reactant and product concentrations remain the same.