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### Question: The equilibrium constant for the reaction: 2 NO (g) <-----> N2 (...)

The equilibrium constant for the reaction: 2 NO (g) <-----> N2 (g) + O2 (g) is 2.60 x 10^-3 at 1100 °C. If 0.820 mole of NO (g) and 0.223 mole each of N2 (g) and O2 (g) are mixed in a 1.00 liter container at 1100 °C, what are the concentrations of NO (g), N2(g), and O2 (g) at equilibrium?

### Answers

Jegannadharo Alla answered this  
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Number of moles of NO = 0.820 mol

Volume of the container = 1.00 L

$$\text{Initial concentration of NO, } [NO] = \frac{0.820 \text{ mol}}{1.00 \text{ L}} = 0.820 \text{ M}$$

Number of moles of N<sub>2</sub> = 0.223 mol

Volume of the container = 1.00 L

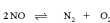
$$\text{Initial concentration of O}_2, [O_2] = \frac{0.223 \text{ mol}}{1.00 \text{ L}} = 0.223 \text{ M}$$

Number of moles of O<sub>2</sub> = 0.223 mol

Volume of the container = 1.00 L

$$\text{Initial concentration of N}_2, [N_2] = \frac{0.223 \text{ mol}}{1.00 \text{ L}} = 0.223 \text{ M}$$

The equilibrium constant for the reaction: 2 NO (g) <====> N<sub>2</sub> (g) + O<sub>2</sub> (g)



Initial concentrations (M): 0.820      0.223      0.223

Change (M):                      -2x                      +x                      +x

Equilibrium (M):                0.820 - 2x    0.223 + x    0.223 + x

Equilibrium constant K<sub>c</sub> = 2.63 × 10<sup>-3</sup>

$$\text{Equilibrium constant } K_c = \frac{[N_2][O_2]}{[NO]^2}$$

$$2.63 \times 10^{-3} = \frac{(0.223 + x)(0.223 + x)}{(0.820 - 2x)^2}$$

By solving x, we get x = -0.164

$$\begin{aligned} \text{Equilibrium concentration of N}_2, [N_2] &= (0.223 + x) \text{ M} \\ &= 0.223 - 0.164 \\ &= 0.059 \end{aligned}$$

$$\begin{aligned} \text{Equilibrium concentration of O}_2, [O_2] &= (0.223 + x) \text{ M} \\ &= 0.223 - 0.164 \\ &= 0.059 \end{aligned}$$

$$\begin{aligned} \text{Equilibrium concentration of NO, } [NO] &= (0.82 - 2(-0.164)) \text{ M} \\ &= 1.15 \text{ M} \end{aligned}$$

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AVREDDY answered this  
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2 NO (g) <-----> N2 (g) + O2 (g)

Kc = 2.60 x 10^-3

initially

[NO] = 0.82 M    [N2] = 0.223 M

[O2] = 0.223 M

Reaction quotient q = [N2][O2]/[NO]^2

$$\begin{aligned} &= (0.223 \cdot 0.223) / (0.82^2) \\ &= 0.07396 \end{aligned}$$

If q > k, backward reaction is favourable (reactants favoured)

so that, at equilibrium

[NO] = 0.82 - 2x

[N2] = 0.223 + x    [O2] = 0.223 + x

Kc = [N2][O2]/[NO]^2

$$(2.6 \cdot 10^{-3}) = (0.223 + x)^2 / (0.82 - 2x)^2$$

x = 0.1644 M

so that [NO] = 0.82 - 0.1644 = 0.9844 M

[N2] = 0.223 - 0.1644 = 0.0596 M

[O2] = 0.223 - 0.1644 = 0.0596 M

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### Practice with similar questions

Q: The equilibrium constant for the reaction: 2 NO(g) + O2(g) = N2(g) is 2.89 x 10^-9 at 900 °C. If 0.821 mole of NO(g) and 0.219 mole each of N2(g) and O2(g) are mixed in a 1.00 liter container at 900 °C, what are the concentrations of NO(g), N2(g), and O2(g) at equilibrium? (10 points) Answer: [NO] = [N] = [O2] =

A: [See answer](#)

### Up next for you in Chemistry

If a dried aspirin sample is contaminated with a trace amount of acetaminophen, what effect could this have on the observed melting point?

[See answer](#)

A student performed the reaction of this experiment (preparation of aspirin) using a water bath at 90 degrees C instead of 50 degrees C. The final product was tested for the presence of phenols with ferric chloride. The test was

[See answer](#)

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### Questions viewed by other students

Q: The equilibrium constant for the reaction: 2 NO (g) + N2 (g) + O2 (g) is K = 2.60 times 10^-3 at 1100 degree C. If 0.820 mole of NO (g) and 0.223 moles each of N2 (g) and O2 (g) are mixed in a 1.00 liter container at 1100 degree C, what are the concentrations of NO (g), N2(g), and O2 (g) at equilibrium?

A: [See answer](#)

Q: A mixture of 1.16 mole of A, 1.35 mole of B and 0.641 mole of C is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium, the number of moles of A is 1.95. Calculate the equilibrium constant, Kc, for the reaction: 2 A (g) <-----> 2 B (g) + C (g).

A: [See answer](#) 👍 100% (3 ratings)

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