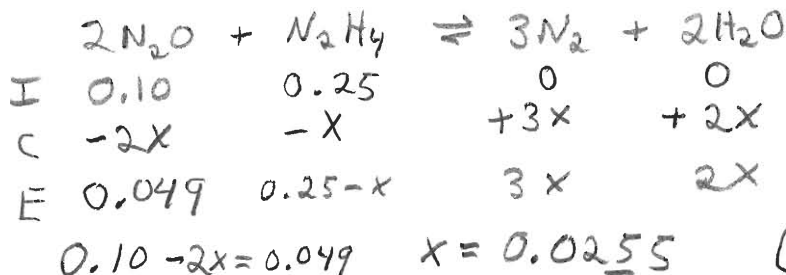


1. The following reaction is investigated (assume an ideal gas mixture):

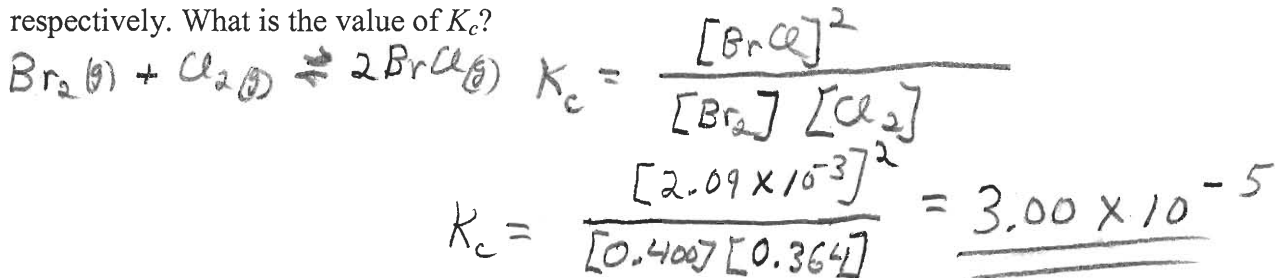


Initially there are 0.10 mol of N_2O and 0.25 mol of N_2H_4 , in a 10.0-L container. If there are 0.049 mol of N_2O at equilibrium, how many moles of N_2 are present at equilibrium?

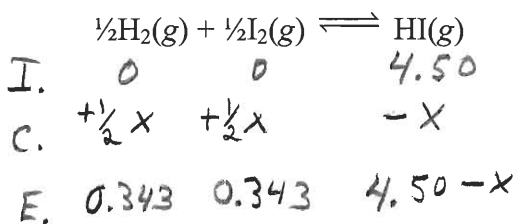


$3x = 0.0765 \text{ moles } \text{N}_2$

2. For the reaction $\text{Br}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{BrCl}(g)$, at equilibrium, it is found that the concentrations of Br_2 , Cl_2 , and BrCl are 0.400 M, 0.364 M, and 2.09×10^{-3} M, respectively. What is the value of K_c ?



3. A 4.50-mol sample of HI is placed in a 1.00-L vessel at 460°C, and the reaction system is allowed to come to equilibrium. The HI partially decomposes, forming 0.343 mol H_2 and 0.343 mol I_2 at equilibrium. What is the equilibrium constant K_c for the following reaction at 460°C?

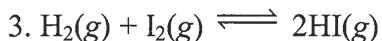
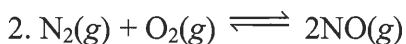
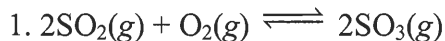


$$K_c = \frac{[3.814]}{[0.343]^{\frac{1}{2}} [0.343]^{\frac{1}{2}}}$$

$K_c = 11.1$

$\frac{1}{2}x = 0.343 \quad x = 0.686 \quad 4.50 - 0.686 = 3.314$

4. For which of the following reactions are the numerical values of K_p and K_c the same?



- A) 1 only
 C) 1 and 2 only
 E) 1, 2, and 3

Δn of gases must be 0

- B) 2 only
D) 2 and 3 only

5. At 298 K, the value of K_c for the reaction $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{HBr}(\text{g})$ is 2.0×10^{19} .

What is K_c for $\text{HBr}(\text{g}) \rightleftharpoons \frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{Br}_2(\text{g})$?

- A) -2.0×10^{19}
- B) 5.0×10^{-20}
- C) 2.2×10^{-10}**
- D) 4.0×10^{-38}
- E) 1.0×10^{19}

$$\frac{[\text{HBr}]^2}{[\text{H}_2][\text{Br}_2]} = 2.0 \times 10^{19}$$

$$\frac{[\text{H}_2]^{\frac{1}{2}}[\text{Br}_2]^{\frac{1}{2}}}{[\text{HBr}]} = ? \quad \frac{1}{\sqrt{2.0 \times 10^{19}}} = 2.2 \times 10^{-10}$$

6. For which of the following reactions will the reactant experience the largest degree of decomposition upon reaching equilibrium at 500 K?

- A) $2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}); K_p = 1.3 \times 10^{-5}$
- B) $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}); K_p = 1.7 \times 10^{-2}$**
- C) $2\text{NO}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{O}_2(\text{g}); K_p = 5.9 \times 10^{-5}$
- D) $2\text{NOF}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{F}_2(\text{g}); K_p = 1.2 \times 10^{-26}$
- E) $2\text{NO}_2\text{F}(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) + \text{F}_2(\text{g}); K_p = 6.6 \times 10^{-22}$

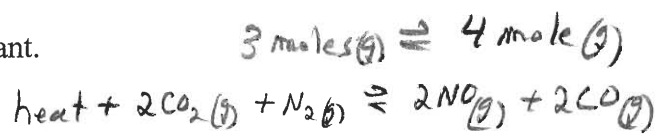
7. $\text{CS}_2(\text{g}) + 3\text{Cl}_2(\text{g}) \rightleftharpoons \text{CCl}_4(\text{g}) + \text{S}_2\text{Cl}_2(\text{g})$

At a given temperature, the reaction above is at equilibrium when $[\text{CS}_2] = 0.050 \text{ M}$, $[\text{Cl}_2] = 0.25 \text{ M}$, $[\text{CCl}_4] = 0.15 \text{ M}$, and $[\text{S}_2\text{Cl}_2] = 0.35 \text{ M}$. What will be the direction of the reaction when the reactants and products have the following concentrations: $\text{CS}_2 = 0.14 \text{ M}$, $\text{Cl}_2 = 0.22 \text{ M}$, $\text{CCl}_4 = 0.32 \text{ M}$, and $\text{S}_2\text{Cl}_2 = 0.33 \text{ M}$?

- A) to the right
- B) to the left**
- C) no change
- D) cannot predict unless we know the temperature
- E) cannot predict unless we know whether the reaction is endothermic or exothermic

8. For the endothermic reaction $2\text{CO}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + 2\text{CO}(\text{g})$, the conditions that favor maximum conversion of the reactants to products are

- A) high temperature and high pressure.
- B) high temperature, pressure being unimportant.
- C) high temperature and low pressure.**
- D) low temperature and high pressure.
- E) low temperature and low pressure.



$$K_c = \frac{[0.15][0.35]}{[0.050][0.25]^3} = 67.2$$

$$Q_c = \frac{[0.32][0.33]}{[0.14][0.22]^3} = 70.8$$

Page 2 Since $Q_c > K_c$ \leftarrow (more react) (less prod)