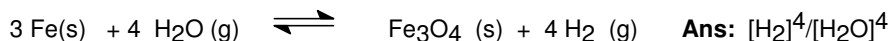


WORKSHEET: CHEMICAL EQUILIBRIUMName _____
Last First**FOR ALL EQUILIBRIUM PROBLEMS, YOU MUST:**

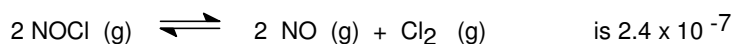
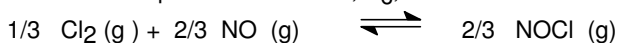
- 1) Write all equilibrium equations
- 2) Write all equilibrium concentrations
- 3) Write all equilibrium expressions

SET A:

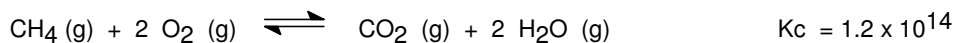
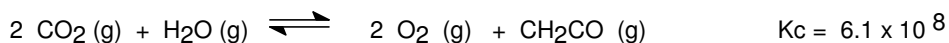
1. a) What is the equilibrium constant expression for the reaction:



- b) The equilibrium constant,
- K_C
- , for the reaction:

What is the equilibrium constant, K_C , for the reaction:**Ans:** 1.6×10^2

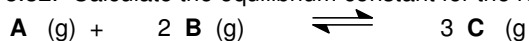
- c) Given the following equilibrium equations and their corresponding equilibrium constants:

Find K_C for the reaction: $\text{CH}_4 \text{ (g)} + \text{CO}_2 \text{ (g)} \rightleftharpoons \text{CH}_2\text{CO (g)} + \text{H}_2\text{O (g)}$

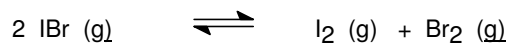
Setup:

Ans: 7.3×10^{22}

2. A mixture of 9.22 moles of
- A**
- , 10.11 moles of
- B**
- , and 27.83 moles of
- C**
- is placed in a one-liter container at a certain temperature. The reaction is allowed to reach equilibrium. At equilibrium the number of moles of
- B**
- is 18.32. Calculate the equilibrium constant for the reaction:

**Ans:** 0.832

3. a. At a certain temperature, K_C is 4.13×10^{-2} for the equilibrium:



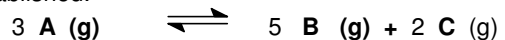
Assume that equilibrium is established at the above temperature by adding only IBr (g) to the reaction flask. What are the concentrations of I_2 (g) and Br_2 (g) in equilibrium with 0.0124 moles/liter of IBr(g) ?
Setup:

Ans: 2.52×10^{-3} M

- b. What was the **initial** concentration of IBr before equilibrium was established?
Setup:

Ans: 0.0174

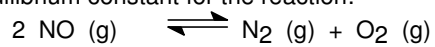
4. 0.924 mole of **A** (g) is placed in a 1.00 liter container at 700°C , where it is 38.8 % dissociated when equilibrium was established.



What is the value of the equilibrium constant, K_C , at the same temperature?
Setup:

Ans: 0.0241

5. The equilibrium constant for the reaction:

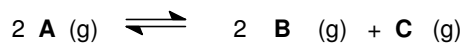


is 2.60×10^{-3} at $1100 \text{ }^\circ\text{C}$. If 0.820 mole of NO (g) and 0.223 mole **each** of N_2 (g) and O_2 (g) are mixed in a 1.00 liter container at $1100 \text{ }^\circ\text{C}$, what are the concentrations of NO (g), N_2 (g), and O_2 (g) at equilibrium?
Setup:

Ans: $[\text{NO}] = 1.15 \text{ M}$
 $[\text{N}_2] = 0.058 \text{ M}$
 $[\text{O}_2] = 0.058 \text{ M}$

SET B :

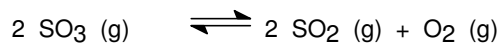
1. A mixture of 1.16 mole of **A**, 1.35 mole of **B** and 0.641 mole of **C** is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium, the number of moles of **A** is 1.95. Calculate the equilibrium constant, K_C , for the reaction:



Setup:

Ans: 0.020

2. 0.822 mole of SO_3 (g) is placed in a 1.00 liter container at 600 K. 36.7 % of the SO_3 (g) are decomposed when equilibrium is established.



What is the value of the equilibrium constant, K_C , at the same temperature?

Setup:

Ans: 5.05×10^{-2}

3. For the equilibrium:

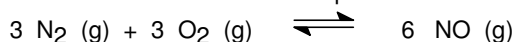


at 205 °C, the equilibrium constant, K_C , is 0.143. If 1.34 moles **each** of Br_2 (g) and Cl_2 (g) are introduced in a container which has a volume of 11.0 liters and allowed to reach equilibrium at 205 °C, what would be the concentrations of Br_2 (g), Cl_2 (g), and BrCl (g) at equilibrium?

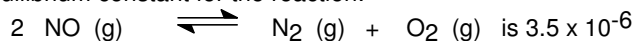
Setup:

Ans: $[\text{Cl}_2] = 0.0529 \text{ M}$
 $[\text{Br}_2] = 0.0529 \text{ M}$
 $[\text{BrCl}] = 0.139 \text{ M}$

4. a. What is the numerical value of the equilibrium constant, K_C , for the reaction:



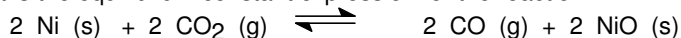
if the equilibrium constant for the reaction:



Setup:

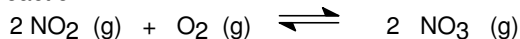
Ans: 2.3×10^{16}

b) What is the equilibrium constant expression for the reaction:



Ans: $K = \frac{[\text{CO}]^2}{[\text{CO}_2]^2}$

5. For the reaction:



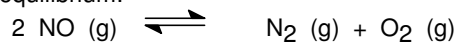
at 923 °C, K_C is 42.5. If 0.0500 mole of NO_2 (g), 0.122 mole of O_2 (g) and 0.300 mole of NO_3 (g) are mixed in a 1.00 liter container at 923 °C, in what direction will the reaction proceed?(Show your calculation to prove that your answer is not a guess.)

Setup:

Answer: $Q_C > K_C$ (The reaction will proceed spontaneously to the left)

SET C :

1. For the equilibrium:

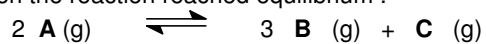


at 300 K, the equilibrium constant, K_C , is 0.185. If 1.45 moles **each** of N_2 (g) and O_2 (g) are introduced in a container that has a volume of 6.00 liters and allowed to reach equilibrium at 300 K, what are the concentrations of N_2 (g), O_2 (g),and NO (g) at equilibrium?

Setup:

Ans: $[\text{N}_2]= 0.112 \text{ M}$
 $[\text{O}_2]= 0.112 \text{ M}$
 $[\text{NO}]= 0.260 \text{ M}$

2. 1.87 mole of **A** (g) are placed in a 1.00 liter container at 700 K. 43.3 % of **A** (g) had decomposed when the reaction reached equilibrium .



What is the value of the equilibrium constant, K_C , at the same temperature?

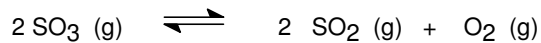
Setup:

Ans: 0.639

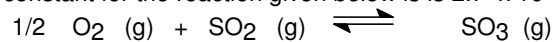
3. a. What is the equilibrium constant expression for the reaction:
 $2 \text{Fe (s)} + 2 \text{NO}_2 \text{(g)} \rightleftharpoons 2 \text{NO (g)} + 2 \text{FeO (s)}$

Ans: $K = \frac{[\text{NO}]^2}{[\text{NO}_2]^2}$

- b. What is the numerical value of the equilibrium constant, K_C , for the reaction:



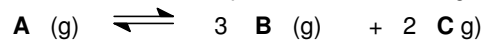
if the equilibrium constant for the reaction given below is 2.7×10^{-3} ?



Setup:

Ans: 1.4×10^5

4. A mixture of 3.31 moles of **A**, 4.33 moles of **B**, and 5.95 moles 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 **B** is 6.16 . Calculate the equilibrium constant, K_C , for the reaction:



Setup:

Ans: 4.5×10^3

5. For the reaction:



at 900 °C, K_c is 1.40×10^{-3} . If 0.780 mole of **A** (g) and 0.244 mole **each** of **B** (g) and **C** (g) are mixed in a 1.00 liter container at 900 °C, what are the concentrations of **A**, **B**, and **C** at equilibrium?

Setup:

Ans: [A]= 1.180 M
[B]= 0.044 M
[C]=0.044 M