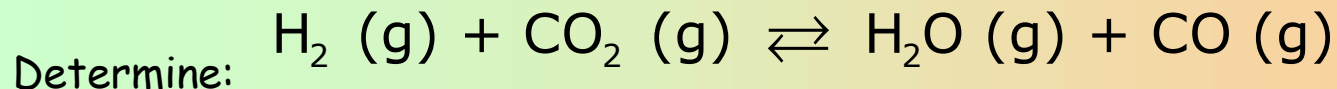


Chemical Equilibrium: exercises

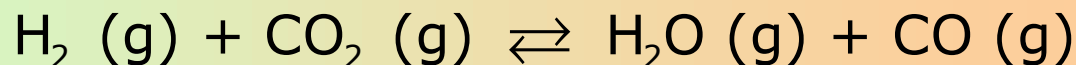
1. The equilibrium constant for this reaction (at 100 °C) is $K_{eq} = 0.772$:



Determine:

a) the concentrations at equilibrium, if the volume of the container is 10 L and the initial composition is 2 moles of H_2 and 2 moles of CO_2 .

The table of number of moles and concentrations:



n initial	2	2	0	0
n change	-x	-x	x	x
n equilibrium	2-x	2-x	x	x
[] equilibrium	(2-x)/10	(2-x)/10	x/10	x/10

Both equilibrium constants (K_p and K_c) are equal:

$$K_p = K_c * (RT)^{\Delta n} \xrightarrow{\Delta n = 0} K_p = K_c = 0.772$$

The value of x is:

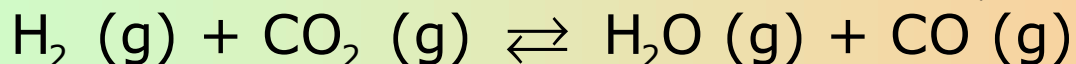
$$K_c = 0.772 = \frac{x^2}{(2-x)^2} \rightarrow 0.88 = \frac{x}{2-x} \rightarrow x = 0.94$$

The concentrations at equilibrium:

$$\begin{aligned} [\text{H}_2\text{O}] &= [\text{CO}] = x/10 = 0.094 \text{ M} \\ [\text{H}_2] &= [\text{CO}_2] = (2-x)/10 = 0.106 \text{ M} \end{aligned}$$

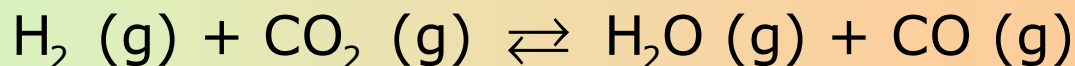
Chemical Equilibrium: exercises

1. The equilibrium constant for this reaction (at 100 °C) is $K_{eq} = 0.772$:



b) the new concentrations at equilibrium if 0.5 mol of H_2 are added after reaching the first equilibrium

The new table of number of moles and concentrations:



n initial	1.56	1.06	0.94	0.94
n change	-x	-x	x	x
n equilibrium	1.56-x	1.06-x	0.94+x	0.94x
[] equilibrium	(1.56-x)/10	(1.06-x)/10	(0.94+x)/10	(0.94+x)/10

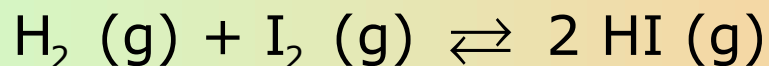
The value of x is: $K_c = 0.772 = \frac{(0.94+x)^2}{(1.56-x)(1.06-x)} \rightarrow x = 0.11$

The new concentrations at equilibrium:

$$\begin{aligned} [\text{H}_2\text{O}] &= [\text{CO}] = (0.94 + x)/10 = 0.105 \text{ M} \\ [\text{CO}_2] &= (1.06 - x)/10 = 0.095 \text{ M} \\ [\text{H}_2] &= (1.56 - x)/10 = 0.145 \text{ M} \end{aligned}$$

Chemical Equilibrium: exercises

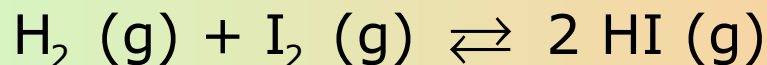
2. The equilibrium constant for this reaction (at 425 °C) is $K_{eq} = 54.8$:



Determine:

a) the concentrations at equilibrium if 0.6 moles of HI are heated in a 500 mL- flask.

The table of number of moles and concentrations:



n initial	0	0	0.6
n change	x	x	-2x
n equilibrium	x	x	0.6-2x
[] equilibrium	2x	2x	2 (0.6-2x)

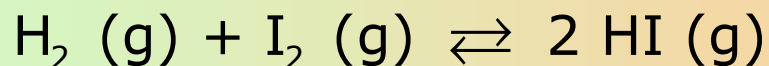
The value of x is: $K_p = K_c = 54.8 = \frac{(0.6-2x)^2}{x^2} \rightarrow 7.4 = \frac{0.6 - 2x}{x} \rightarrow x = 0.064$

The concentrations at equilibrium:

$$\begin{aligned} [\text{HI}] &= 2 (0.6 - 2x) = 0.944 \text{ M} \\ [\text{H}_2] &= [\text{I}_2] = 2x = 0.128 \text{ M} \end{aligned}$$

Chemical Equilibrium: exercises

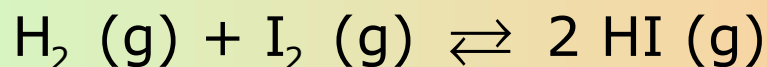
2. The equilibrium constant for this reaction (at 425 °C) is $K_{eq} = 54.8$:



Determine:

b) the new concentrations at equilibrium if 0.1 mol of H_2 are added after reaching the first equilibrium

The table of number of moles and concentrations:



n initial	0.164	0.064	0.472
n change	-x	-x	+2x
n equilibrium	0.164-x	0.064-x	0.472+2x
[] equilibrium	0.328-2x	0.128-2x	0.944+4x

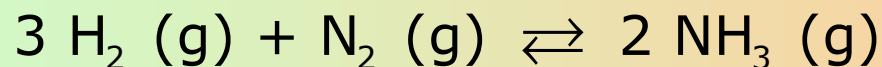
The value of x is: $K_c = 54.8 = \frac{(0.944 + 4x)^2}{(0.328-2x) * (0.128-2x)} \rightarrow x = 0.027$

The new
concentrations at
equilibrium:

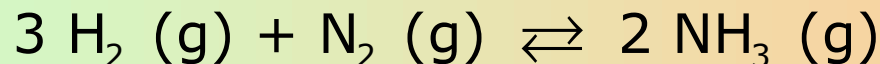
$$\begin{aligned} [\text{HI}] &= 0.944 + 4x = 1.052 \text{ M} \\ [\text{I}_2] &= 0.128 - 2x = 0.074 \text{ M} \\ [\text{H}_2] &= 0.328 - 2x = 0.274 \text{ M} \end{aligned}$$

Chemical Equilibrium: exercises

3. A 0.5 L-flask, at 725 °C, is charged with 0.4 moles of ammonia. When equilibrium is reached 0.1 mol of ammonia remain in the flask. Determine the equilibrium constant for this reaction (in terms of concentration)



The table of number of moles and concentrations:



n initial	0	0	0.4
n change	3x	x	-2x
n equilibrium	3x	x	0.4-2x 0.1 mol
[] equilibrium	0.9 M	0.3 M	0.2 M

$$0.4 - 2x = 0.1$$

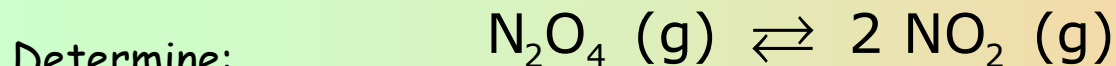
$$x = 0.15$$

The equilibrium constant:

$$K_c = \frac{[\text{NH}_3]^2}{[\text{H}_2]^3 [\text{N}_2]} = \frac{(0.2)^2}{(0.9)^3 (0.3)} = 0.183$$

Chemical Equilibrium: exercises

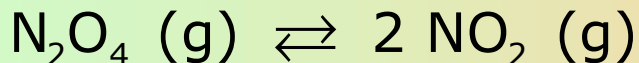
4. At 150 °C, 69% of dinitrogen tetraoxide is dissociated when pressure is 4 atm.



Determine:

a) K_p b) the pressure for which equilibrium is reached when the percent dissociation is 80 %

a) The table of number of moles and concentrations:



p initial	p_0	0
p change	$- p_0 \alpha = -0.69 p_0$	$2 p_0 \alpha = 1.38 p_0$
p equilib.	$p_0 (1-\alpha) = 0.31 p_0$ 0,7338 atm	$2 p_0 \alpha = 1.38 p_0$ 3,2665 atm

p total
p_0
$0.69 p_0$
$1.69 p_0$
4 atm

$$p_0 = \frac{4 \text{ atm}}{1.69} = 2.367 \text{ atm}$$

The equilibrium constant: $K_p = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} = \frac{4 \alpha^2}{1-\alpha} P_0 = \frac{(3.2665)^2}{(0.7338)} = 14.54$

$$K_p = 14.54$$

b) The table of number of moles and concentrations:

p initial	p_0	0
p change	$- p_0 \alpha = -0.80 p_0$	$2 p_0 \alpha = 1.60 p_0$
p equilib.	$p_0 (1-\alpha) = 0.20 p_0$	$2 p_0 \alpha = 1.60 p_0$

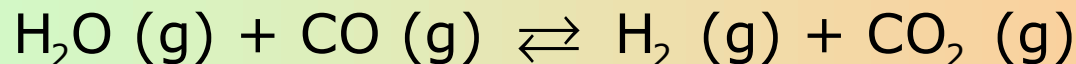
p total
p_0
$0.80 p_0$
$1.80 p_0$

$$P_T = 2.0447 \text{ atm}$$

The total pressure: $K_p = 14.54 = \frac{4 \alpha^2}{1-\alpha} P_0 = \frac{4 (0.8)^2}{1-0.8} P_0 \rightarrow P_T = 1.8 * P_0 = 2.0447 \text{ atm}$

Chemical Equilibrium: exercises

5. The equilibrium constant for this reaction (at 986 °C) is $K_{eq} = 0.63$:



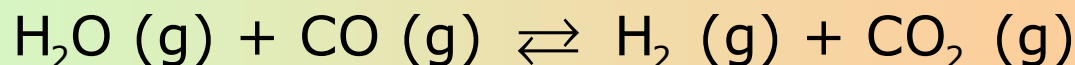
A flask is charged with 1 mol of water vapour and 3 moles of CO and equilibrium is reached at 986 °C and 2 atm.

Determine:

a) the number of moles of hydrogen at equilibrium

Both equilibrium constants are equal: $K_p = K_c * (RT)^{\Delta n} \xrightarrow{\Delta n = 0} K_p = K_c = 0.63$

The table of number of moles and concentrations:

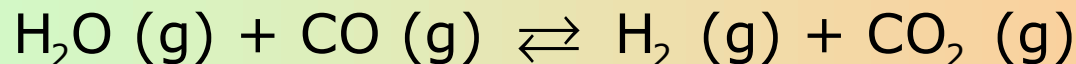


n initial	1	3	0	0
n change	-x	-x	x	x
n equilibrium	1-x	3-x	x	x
[] equilibrium	(1-x)/V	(3-x)/V	x/V	x/V

The value of x and the # of moles of H₂ are: $K_c = 0.63 = \frac{x^2}{(1-x)(3-x)} \rightarrow x = 0.68 = n(\text{H}_2)$

Chemical Equilibrium: exercises

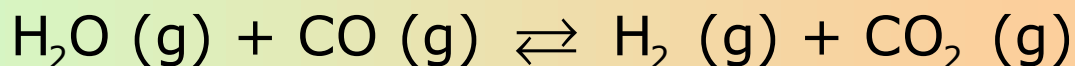
5. The equilibrium constant for this reaction (at 986 °C) is $K_{eq} = 0.63$:



A flask is charged with 1 mol of water vapour and 3 moles of CO and equilibrium is reached at 986 °C and 2 atm.

Determine:

b) the partial pressure of all components at equilibrium



n initial	1	3	0	0
n change	-x	-x	x	x
n equilibrium	1-x	3-x	x	x
[] equilibrium	(1-x)/V	(3-x)/V	x/V	x/V

The partial pressures are:

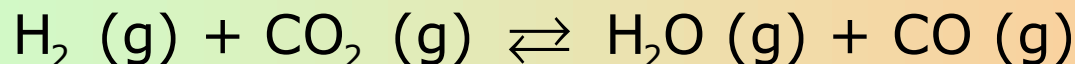
$$P_{\text{H}_2\text{O}} = \frac{n(\text{H}_2\text{O})}{n(\text{total})} * P_T = \frac{1-x}{4} * 2 \text{ atm} \rightarrow P_{\text{H}_2\text{O}} = 0.16 \text{ atm}$$

$$P_{\text{H}_2} = P_{\text{CO}_2} = \frac{x}{4} * 2 \text{ atm} \rightarrow P_{\text{H}_2} = P_{\text{CO}_2} = 0.34 \text{ atm}$$

$$P_{\text{CO}} = \frac{3-x}{4} * 2 \text{ atm} \rightarrow P_{\text{CO}} = 1.16 \text{ atm}$$

Chemical Equilibrium: exercises

6. A 10 L-container is charged with 0.61 moles of CO_2 and 0.39 moles of H_2 and the mixture is heated to $1250\text{ }^\circ\text{C}$. When the equilibrium is reached 0.35 moles of CO_2 are found inside. Determine:



- the composition of the other gases
- the values of K_c and K_p at that temperature

The table of number of

moles and concentrations: $\text{H}_2 (\text{g}) + \text{CO}_2 (\text{g}) \rightleftharpoons \text{H}_2\text{O} (\text{g}) + \text{CO} (\text{g})$

n initial	0.39	0.61	0	0
n change	-x	-x	x	x
n equilibrium	0.39-x	0.61-x = 0.35	x	x
[] equilibrium	0.013 M	0.035 M	0.026 M	0.026 M

$$x = 0.26$$

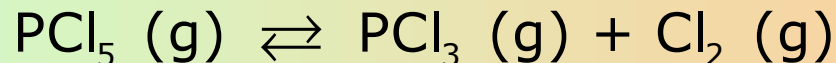
The equilibrium constants:

$$K_c = \frac{[\text{H}_2\text{O}] [\text{CO}]}{[\text{H}_2] [\text{CO}_2]} = \frac{(0.026)^2}{(0.013) (0.035)} = 1.486$$

$$K_p = K_c * (RT)^{\Delta n} \rightarrow K_p = K_c = 1.486$$

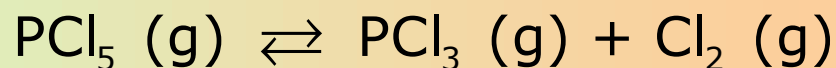
Chemical Equilibrium: exercises

7. In a 10 L-container the composition of the mixture at equilibrium (at 80 °C) is: 0.3 mol PCl_5 , 0.7 mol PCl_3 , and 0.7 mol Cl_2 . Determine:



a) the equilibrium constants K_c and K_p

The table of number of moles and concentrations:



n equilibrium	0.3	0.7	0.7
[] equilibrium	0.03 M	0.07 M	0.07 M

The equilibrium constants:

$$K_c = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]} = \frac{(0.07)^2}{(0.03)} = 0.163$$

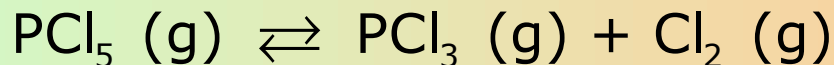
$$K_p = K_c * (RT)^{\Delta n} \rightarrow$$

$$K_p = K_c * (RT)^{\Delta n} = 0.163 * (0.082 * 353)$$

$$K_p = 4.718$$

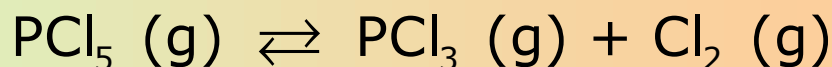
Chemical Equilibrium: exercises

7. In a 10 L-container the composition of the mixture at equilibrium (at 80 °C) is: 0.3 mol PCl_5 , 0.7 mol PCl_3 , and 0.7 mol Cl_2 . Determine:



b) if 0.3 mol of chlorine gas is added after equilibrium has been reached, the number of moles of each component when equilibrium is reached again.

The table of number of moles and concentrations:



n initial	0.3	0.7	1
n change	x	-x	-x
n equilibrium	0.3+x	0.7-x	1-x
[] equilibrium	(0.3+x)/10	(0.7-x)/10	(1-x)/10

The new concentrations at equilibrium:

$$K_c = 0.163 = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]} = \frac{(0.7 - x) (1-x)}{10 (0.3 + x)}$$

$$x^2 - 3.33 x + 0.211 = 0 \rightarrow x = 0.065$$

$$n (\text{PCl}_5) = 0.3 + x = 0.365 \text{ mol}$$

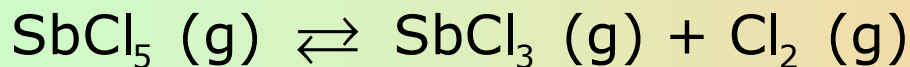
$$n (\text{PCl}_3) = 0.7 - x = 0.635 \text{ mol}$$

$$n (\text{Cl}_2) = 1 - x = 0.935 \text{ mol}$$

Chemical Equilibrium: exercises

8. At 185 °C and 1 atm antimony pentachloride (SbCl_5) dissociates partially. Its percent dissociation is 30% and the products of this dissociation are antimony trichloride (SbCl_3) and chlorine. Determine:
a) the dissociation equation for 1 mole of antimony pentachloride and the equilibrium constants at that temperature

The table of number of moles and concentrations:



p initial	p_0	0	0
p change	$-p_0\alpha$ $-0.3p_0$	$p_0\alpha$ $0.3p_0$	$p_0\alpha$ $0.3p_0$
p equilibrium	$p_0(1-\alpha)$	$p_0\alpha$	$p_0\alpha$

p total
p_0
$0.3 p_0$
$1.3 p_0$
1 atm

$$p_0 = 0.769 \text{ atm}$$

The equilibrium constants:

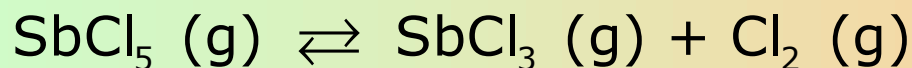
$$K_p = \frac{P_{\text{SbCl}_3} P_{\text{Cl}_2}}{P_{\text{SbCl}_5}} = \frac{(0.3)^2 p_0}{0.7} = 0.0989$$

$$K_c = \frac{K_p}{RT} = \frac{0.0989}{(0.082 \cdot 458)} = 2.63 \cdot 10^{-3}$$

Chemical Equilibrium: exercises

8. At 185 °C and 1 atm antimony pentachloride (SbCl_5) dissociates partially. Its percent dissociation is 30% and the products of this dissociation are antimony trichloride (SbCl_3) and chlorine. Determine:
b) the pressure needed to increase the percent dissociation to 60%

The table of number of moles and concentrations:



p initial	p_0	0	0
p change	$-p_0\alpha$ $-0.6p_0$	$p_0\alpha$ $0.6p_0$	$p_0\alpha$ $0.6p_0$
p equilibrium	$p_0(1-\alpha)$	$p_0\alpha$	$p_0\alpha$

p total
p_0
$0.6 p_0$
$1.6 p_0$

In the case of 60%:

$$n_T = n(\text{SbCl}_5) + n(\text{SbCl}_3) + n(\text{Cl}_2) = n_0 (1-\alpha) + n_0\alpha + n_0\alpha = n_0 (1+\alpha)$$

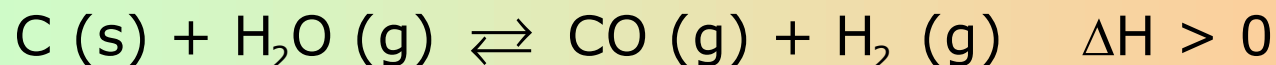
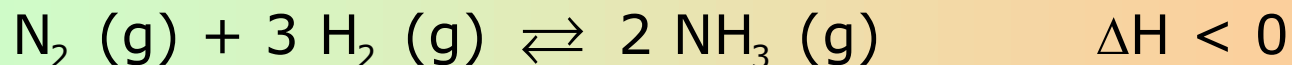
$$P_{\text{SbCl}_5} = \frac{1-\alpha}{1+\alpha} P_T = 0.25 P_T; \quad P_{\text{SbCl}_3} = P_{\text{Cl}_2} = \frac{\alpha}{1+\alpha} P_T = 0.375 P_T$$

$$K_p = 0.0989 = \frac{P_{\text{SbCl}_3} P_{\text{Cl}_2}}{P_{\text{SbCl}_5}} = \frac{(0.375)^2 p_T}{0.25}$$

$$p_T = 0.176 \text{ atm}$$

Chemical Equilibrium: exercises

9. Determine what happens in these reactions when equilibrium is disturbed as follows



- a) increase in pressure b) increase in temperature
c) increase in the amount of hydrogen d) addition of a catalyst

a) Increase in pressure:
reaction #1 ... shifts to the right
reaction #2 ... shifts to the left
(carbon is solid)

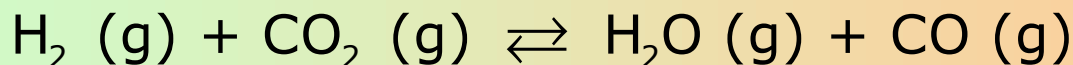
b) Increase in temperature:
reaction #1 ... shifts to the left
reaction #2 ... shifts to the right

c) Increase in amount of hydrogen:
reaction #1 ... shifts to the right
reaction #2 ... shifts to the left

d) A catalyst:
reaction #1 ... no effect
reaction #2 ... no effect

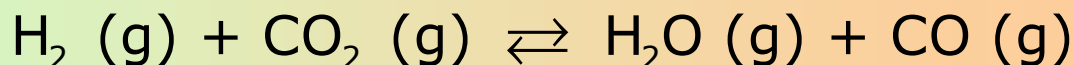
Chemical Equilibrium: exercises

10. A 10 L-container is charged with 2 moles of H_2 and 2 moles of CO_2 at $100^\circ C$. The equilibrium constant in terms of concentration is 0.772. Determine:



- the concentrations at equilibrium of all components
- K_p and the total pressure at equilibrium

The table of number of moles and concentrations:



n initial	2	2	0	0
n change	-x	-x	x	x
n equilibrium	2-x	2-x	x	x
[] equilibrium	(2-x)/10	(2-x)/10	x/10	x/10

Both equilibrium constants are equal: $K_p = K_c * (RT)^{\Delta n} \xrightarrow{\Delta n = 0} K_p = K_c = 0.772$

The value of x is: $K_c = 0.772 = \frac{x^2}{(2-x)^2} \rightarrow 0.88 = \frac{x}{2-x} \rightarrow x = 0.94$

The concentrations at equilibrium:

$$[H_2O] = [CO] = x/10 = 0.094 \text{ M}$$

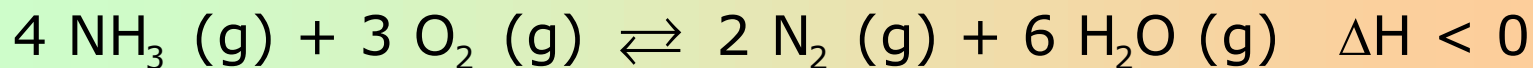
$$[H_2] = [CO_2] = (2-x)/10 = 0.106 \text{ M}$$

The total pressure:

$$P = c R T = 0.4 \frac{\text{mol}}{\text{L}} * 0.082 \frac{\text{atm.L}}{\text{K.mol}} * 373 \text{ K} = 12.23 \text{ atm}$$

Chemical Equilibrium: exercises

11. Determine what happens in this reactions when equilibrium is disturbed as follows



- a) decrease in pressure b) increase in temperature
c) elimination of water d) addition of a catalyst

a) decrease in pressure:
reaction ... shifts to the right

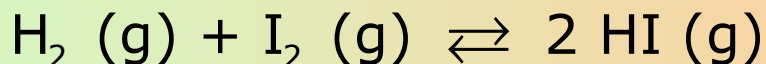
b) Increase in temperature:
reaction ... shifts to the left

c) Elimination of water:
reaction ... shifts to the right

d) A catalyst:
reaction ... no effect

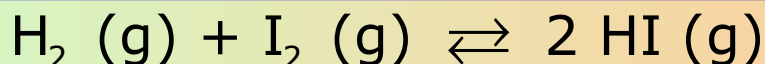
Chemical Equilibrium: exercises

12. A 100 L-container is charged with 2 moles of H_2 and 4 moles of I_2 and the mixture is heated to $400\text{ }^\circ\text{C}$. At that temperature the equilibrium constant is $K_c=55$. Determine:



a) the number of moles of all components at equilibrium

b) if the container is opened and the external pressure is 1 atm, the direction of flow of gases



n initial	2	4	0
n change	-x	-x	2x
n equilibrium	2-x	4-x	2x
[] equilibrium	(2-x)/100	(4-x)/100	2x/100

The value of x is: $K_c = 55 = \frac{4 x^2}{(2-x)(4-x)} \rightarrow 51 x^2 - 330 x + 440 = 0 \rightarrow x = 1.88$

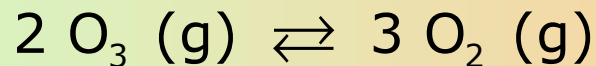
The number of moles: $n(H_2) = 2 - x = 0.12 \text{ mol}$; $n(I_2) = 4 - x = 2.12 \text{ mol}$
 $n(HI) = 2 x = 3.76 \text{ mol}$

The total pressure: $p_T = c R T = 0.06 \frac{\text{mol}}{\text{L}} * 0.082 \frac{\text{atm.L}}{\text{K.mol}} * 673 \text{ K} = 3.31 \text{ atm}$

The gases will get out of the vessel

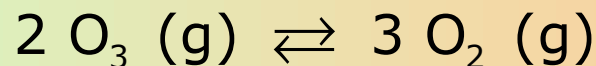
Chemical Equilibrium: exercises

13. At 1127 °C the pressure of a mixture of ozone and oxygen at equilibrium is 18.1 atm. If the percent dissociation of ozone is 97 %, determine



a) the value of K_p

b) when the equilibrium is disturbed by heating, the concentration of ozone increases and the concentration of oxygen decreases. ¿Is the process exothermic or endothermic?



p initial	n_0	0
p change	$-n_0 \alpha = -0.97 n_0$	$1.5 n_0 \alpha = 1.455 n_0$
p equilib.	$n_0 (1-\alpha) = 0.03 n_0$	$1.5 n_0 \alpha = 1.455 n_0$

$$n_{\text{tot}} = 1.485 n_0$$

$$p_{\text{tot}} = 18.1 \text{ atm}$$

$$\text{Molar fractions: } x(\text{O}_2) = \frac{n(\text{O}_2)}{n_{\text{tot}}} = \frac{1.455 n_0}{1.485 n_0} = 0.9798; \quad x(\text{O}_3) = 0.0202$$

$$\text{Partial pressures: } P(\text{O}_2) = x(\text{O}_2) \cdot P_{\text{total}} = 17.734 \text{ atm}$$

$$P(\text{O}_3) = x(\text{O}_3) \cdot P_{\text{total}} = 0.3656 \text{ atm}$$

$$K_p = \frac{P_{\text{O}_2}^3}{P_{\text{O}_3}^2} = \frac{(17.734)^3}{(0.3656)^2} = 41726$$

The process is exothermic.