

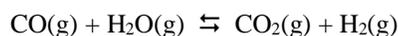
WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NOTE: EXAM 2 is **Wednesday, October 10th**. It will cover Chapter 15, and Sections 16.1, 16.3, 16.4, and 16.5 of Chapter 16. This includes the Chapter 16 ppt slides 1-31, and the Chapter 16 problems 1-10.

NAME _____ Panther ID _____

For problems involving calculations you must show your work for credit.

1) The equilibrium constant for the reaction



is $K_C = 5.10$ at $T = 700. \text{ K}$

A system initially has $[\text{CO}] = 0.0500 \text{ M}$, $[\text{H}_2\text{O}] = 0.1000 \text{ M}$, and $[\text{CO}_2] = 0.0800 \text{ M}$. There is initially no H_2 in the system. What will be the concentrations of H_2O and H_2 in the system when equilibrium is reached?

$$K_C = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} = 5.10$$

We first set up an ICE table

	Initial	Change	Equilibrium
CO_2	0.0800	x	0.0800 + x
H_2	0	x	x
CO	0.0500	- x	0.0500 - x
H_2O	0.1000	- x	0.1000 - x

$$\frac{(0.0800 + x)(x)}{(0.0500 - x)(0.1000 - x)} = 5.10$$

We could try assuming that $x \ll 0.0500$, but since $K_C \sim 1$ this likely will not be a good assumption. So we can proceed as follows.

$$(0.0800 + x)(x) = 5.10(0.0500 - x)(0.1000 - x)$$

$$0.0800x + x^2 = 5.10x^2 - 0.765x + 0.0255$$

$$4.10x^2 - 0.845x + 0.0255 = 0$$

$$x = \frac{0.845 \pm [(-0.845)^2 - 4(4.10)(0.0255)]^{1/2}}{2(4.10)}$$

$$= \frac{0.845 \pm 0.544}{8.20} = 0.169, \underline{0.0367}$$

The underlined root is the only one that gives all positive values for concentration. Therefore, at equilibrium

$$\begin{aligned} [\text{H}_2\text{O}] &= 0.0633 \text{ M} & [\text{CO}] &= 0.0133 \text{ M} \\ [\text{H}_2] &= 0.0367 \text{ M} & [\text{CO}_2] &= 0.1167 \text{ M} \end{aligned}$$

As a check, if we substitute the above into the expression for K_C we get $K_C = \frac{(0.1167)(0.0367)}{(0.0133)(0.0633)} = 5.09$

correct to roundoff error.

2) Consider the following chemical reaction at T = 300. K



A closed system containing all of the above substances is initially at equilibrium. Predict (by circling the correct answer) whether each of the following changes to the system will cause the moles of PCl_3 to increase, to decrease, or to stay the same.

Addition of 0.100 moles of Cl_2 to the system

moles PCl_3 will increase	moles PCl_3 will stay the same	<u>moles PCl_3 will decrease</u>
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Increasing the volume of the system by 4.00 L, while keeping temperature constant

<u>moles PCl_3 will increase</u>	moles PCl_3 will stay the same	moles PCl_3 will decrease
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Increasing the temperature of the system by 25.0 K

<u>moles PCl_3 will increase</u>	moles PCl_3 will stay the same	moles PCl_3 will decrease
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For the first change, adding Cl_2 to the system will lead to the system trying to get rid of Cl_2 . For the second change, increasing the volume decreases the total pressure of the system, and so the system will act in a way that increases total pressure. For the third change, since the reaction is endothermic heat is (in a sense) a reactant, so if we increase temperature the system will try to remove heat by shifting to the right.

3) Consider the following chemical reaction: $2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$

a) Write the expression for K (the thermodynamic equilibrium constant) for the above reaction.

$$K = \frac{(p_{\text{NO}_2})^2}{(p_{\text{NO}})^2 (p_{\text{O}_2})}$$

b) Using the thermochemical data given below (at T = 25.0 °C) find the numerical value for K for the above reaction at T = 25.0 °C.

substance	ΔH°_f (kJ/mol)	ΔG°_f (kJ/mol)	S° (J/mol·K)
NO(g)	90.25	86.55	210.76
NO ₂ (g)	33.18	51.31	240.06
O ₂ (g)	0.00	0.00	205.14

$$\begin{aligned} \Delta G^\circ_{\text{rxn}} &= [2 \Delta G^\circ_f(\text{NO}_2(\text{g}))] - [2 \Delta G^\circ_f(\text{NO}(\text{g})) + \Delta G^\circ_f(\text{O}_2(\text{g}))] \\ &= [2 (51.31)] - [2 (86.55) + (0.00)] = - 70.48 \text{ kJ/mol} \end{aligned}$$

$$\ln K = \frac{\Delta G^\circ_{\text{rxn}}}{RT} = - \frac{(- 70480. \text{ J/mol})}{(8.314 \text{ J/mol}\cdot\text{K}) (298. \text{ K})} = 28.45 \quad K = e^{28.55} = 2.3 \times 10^{12}$$