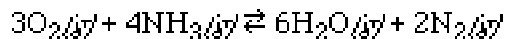


ALL work must be shown to receive full credit. **Due in lecture at 8:30 a.m. on Wednesday, March 13, 2002.**

PS7.1. Given the reaction



Initially (before any reaction occurs) a 1.00 liter reaction vessel at 400 °C contains 0.502 moles of O₂ and 0.791 moles of NH₃ and no water or nitrogen. Consider the following:

- a) If 0.0873 moles of O₂ react, how many moles of NH₃ must react and how many moles of H₂O and N₂ are formed? How many moles of O₂, NH₃, H₂O and N₂ remain after completion of the reaction?

	3O₂	+	4NH₃	⇌	6H₂O	+	2N₂
initial	.502		.791		0		0
change	-.0873		-.116		+.175		+.0582
moles remaining(final)	.415		.675		.175		.0582

Note: the change row was determined using the reaction stoichiometry calculations below,

$$0.0873 \text{ mol O}_2 \left(\frac{4 \text{ mol NH}_3}{3 \text{ mol O}_2} \right) = 0.116 \text{ mol NH}_3 \text{ reacting}$$

$$0.0873 \text{ mol O}_2 \left(\frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} \right) = 0.175 \text{ mol H}_2\text{O formed}$$

$$0.0873 \text{ mol O}_2 \left(\frac{2 \text{ mol N}_2}{3 \text{ mol O}_2} \right) = 0.0582 \text{ mol N}_2 \text{ formed}$$

The moles remaining row is determined by adding the moles reacting or forming (change row) to the initial amount of each species.

- b) If 0.234 moles of NH₃ react, how many moles of O₂ must react and how many moles of H₂O and N₂ are formed? How many moles of O₂, NH₃, H₂O and N₂ remain after completion of the reaction?

	3O₂	+	4NH₃	⇌	6H₂O	+	2N₂
initial	.502		.791		0		0
change	-.176		-.234		+.351		+.117
moles remaining(final)	.326		.557		.351		.117

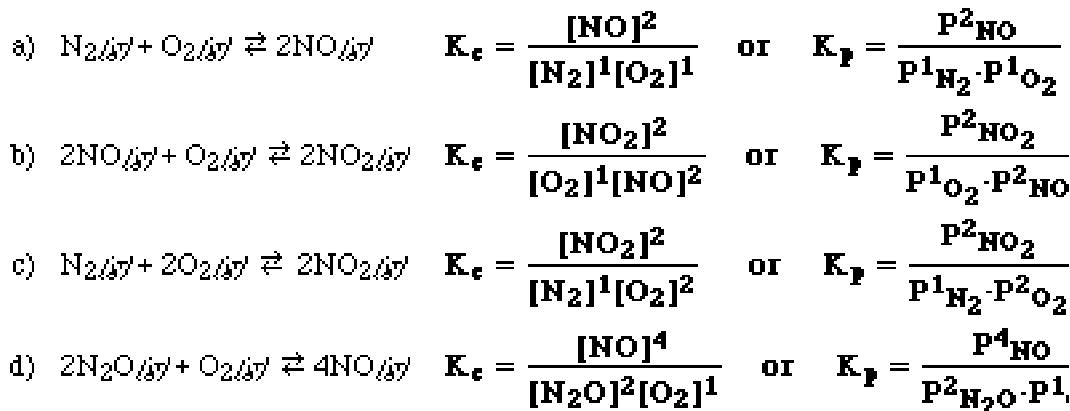
- c) If '3x' moles of O₂ react, how many moles of NH₃ must react and how many moles of H₂O and N₂ are formed (in terms of 'x')? How many moles of O₂, NH₃, H₂O and N₂ remain after completion of the reaction?

	3O₂	+	4NH₃	⇌	6H₂O	+	2N₂
initial	.502		.791		0		0
change	-3x		-4x		+6x		+2x
moles remaining(final)	.502-3x		.791-4x		0+6x		0+2x

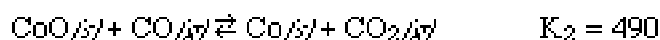
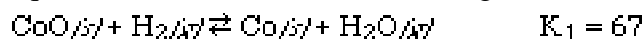
- d) If 0.875 moles of H₂O are formed, how many moles of N₂ are formed and how many moles of O₂ and NH₃ must react? How many moles of O₂, NH₃, H₂O and N₂ remain after completion of the reaction?

	3O₂	+	4NH₃	⇌	6H₂O	+	2N₂
initial	.502		.791		0		0
change	-.438		-.583		+.875		+.292
moles remaining(final)	.064		.208		.875		.292

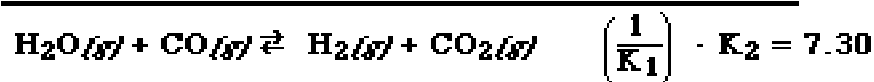
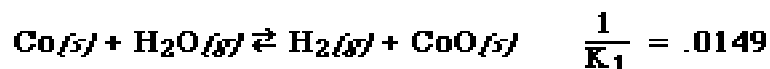
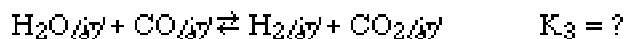
PS7.2. Write the equilibrium expression for each of the following chemical equations;



PS7.3. Equilibrium constants for the following reactions have been determined at 550 °C:



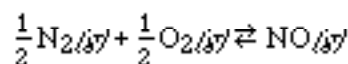
Calculate K_c (at the same temperature) for the commercially important water gas shift reaction



PS7.4. Calculate K_c for the reaction



if K_c for the reaction

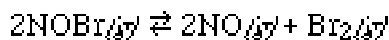


is 1.3×10^4 .

$$K_c = \frac{[\text{NO}]^1}{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}} = 1.3 \times 10^4$$

$$K_c' = \frac{[\text{NO}]^2}{[\text{N}_2]^1[\text{O}_2]^1} = (K_c)^2 = (1.3 \times 10^4)^2 = 1.7 \times 10^8$$

PS7.5. A 1.00 liter container initially holds 0.257 moles of NOBr at a given temperature. The reaction which occurs is:



At equilibrium analysis shows 0.240 moles of NO and 0.120 moles of Br₂.

- a) Which direction did the reaction proceed to establish (reach) equilibrium?

Reaction proceeds from left to right (L → R)

- b) How many moles of NOBr reacted in order to form 0.240 moles of NO and 0.120 moles of Br₂?

If 0.240 mol of NO are formed in the reaction, then;

$$0.240 \text{ mol NO} \left(\frac{2 \text{ mol NOBr}}{2 \text{ mol NO}} \right) = 0.240 \text{ mol NOBr reacted}$$

If 0.120 mol of Br₂ are formed in the reaction, then;

$$0.120 \text{ mol Br}_2 \left(\frac{2 \text{ mol NOBr}}{1 \text{ mol Br}_2} \right) = 0.240 \text{ mol NOBr reacted}$$

- c) How many moles of NOBr remain after equilibrium was established?

$$[\text{NOBr}]_{\text{eq}} = [\text{NOBr}]_0 + [\text{NOBr}]_{\text{reacted}} = 0.257 \text{ M} - 0.240 \text{ M}$$

$$[\text{NOBr}]_{\text{eq}} = 0.017 \text{ M}$$

- d) What is the magnitude of K_c?

$$K_c = \frac{[\text{NO}]^2[\text{Br}_2]}{[\text{NOBr}]^2} = \frac{[0.240]^2[0.120]}{[0.017]^2} = 23.9$$

PS7.6. In a container, the partial pressure of NOCl is initially 0.340 atm at a given temperature. The chemical equation which describes the reaction is:



At equilibrium analysis shows the partial pressure of NO is 0.0916 atm.

- a) Which direction did the reaction proceed to establish (reach) equilibrium?

Reaction proceeds from right to left (R → L)

- b) What is the partial pressure of NOCl which reacted in order for the partial pressure of NO to be 0.0916 atm?

For the partial pressure of NO to be 0.0916 atm, the amount of NOCl which must react is,

$$0.0916 \text{ atm NO} \left(\frac{2 \text{ atm NOCl}}{2 \text{ atm NO}} \right) = 0.0916 \text{ atm NOCl}$$

- c) What is the partial pressure of Cl₂ at equilibrium?

Since the initial partial pressure of Cl₂ is zero, the amount of Cl₂ formed is equal to the amount at equilibrium.

$$0.0916 \text{ atm NO} \left(\frac{1 \text{ atm Cl}_2}{2 \text{ atm NO}} \right) = 0.0458 \text{ atm Cl}_2$$

$$(\text{P}_{\text{Cl}_2})_{\text{eq}} = (\text{P}_{\text{Cl}_2})_0 + (\text{P}_{\text{Cl}_2})_{\text{forming}}$$

$$(\text{P}_{\text{Cl}_2})_{\text{eq}} = 0 + 0.0458 \text{ atm} = 0.0458 \text{ atm}$$

- d) What is the partial pressure of NOCl at equilibrium?

$$(\text{P}_{\text{NOCl}})_{\text{eq}} = (\text{P}_{\text{NOCl}})_0 - (\text{P}_{\text{NOCl}})_{\text{reacted}}$$

$$(\text{P}_{\text{NOCl}})_{\text{eq}} = 0.340 \text{ atm} - 0.0916 \text{ atm} = 0.248 \text{ atm}$$

- e) What is the magnitude of K_p?

$$K_p = \frac{\text{P}_{\text{NOCl}}^2}{\text{P}_{\text{NO}}^2 \cdot \text{P}_{\text{Cl}_2}} = \frac{(0.248)^2}{(0.0916)^2(0.0458)} = 160$$

PS7.7. A 1.00 liter container holds 1.06 moles of H₂ and 1.57 moles of CO at a temperature of 162 °C. At this temperature, the following reaction occurs,



After equilibrium is established, analysis shows 0.200 moles of CH₃OH in the container. Calculate the [CO]_{eq}, [H₂]_{eq} and K_c.

	2H₂(g)	+	CO(g)	⇌	CH₃OH(g)	
Initial	1.06 M		1.57 M		0	let x = [CO] _{reacting}
Change	-2x		-x		+x	
Equilibrium	1.06 - 2x		1.57 - x		0 + x	

$$[\text{CH}_3\text{OH}]_{\text{eq}} = 0.200 \text{ M} = x$$

$$[\text{H}_2]_{\text{eq}} = 1.06 - 2x = 1.06 - 2(0.2 \text{ M}) = 0.660 \text{ M}$$

$$[\text{CO}]_{\text{eq}} = 1.57 - x = 1.57 - (0.2 \text{ M}) = 1.37 \text{ M}$$

$$K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{H}_2]^2[\text{CO}]} = \frac{(0.200)}{(0.660)^2 \cdot (1.37)} = 0.335$$

PS7.8. The following reaction,



occurs at 298K. If 2.00 mol of HI are placed into a 1.00 liter container and permitted to react, at equilibrium it is found that 20.0 % of the HI has decomposed. Calculate K_c and K_p.

	2HI	⇌	H₂	+	I₂
initial	2.00		0		0
change	-2x		+x		+x
equilibrium	1.60		+0.2		+0.2

At equilibrium, 20% of the HI has decomposed; therefore, 0.20(2.0 M) = 0.4 M is the amount of HI that reacts.

$$0.4 \text{ M} = 2x$$

$$0.2 \text{ M} = x$$

Substituting

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{(0.2)(0.2)}{(1.6)^2} = 0.0156$$

$$K_p = K_c = .0156$$

PS7.9 The equation which describes the preparation of ammonia is:



A 3.000 L reaction vessel initially contains 0.3000 moles N_2 and 0.4500 moles H_2 . When the reaction is allowed to attain equilibrium at a given temperature, analysis determines 0.09992 M N_2 . Calculate K_c for the reaction.

	$\text{N}_2(g)$	+	$3\text{H}_2(g)$	\rightleftharpoons	$2\text{NH}_3(g)$	
Initial	0.1000 M		0.1500 M		0	let $x = [\text{N}_2]_{\text{reacting}}$
Change	-x		-3x		+2x	
Equilibrium	0.100 - x		0.150 - 3x		0 + 2x	

$$[\text{N}_2]_{\text{eq}} = 0.1000 \text{ M} - x = 0.09992 \text{ M}$$

$$x = 0.1000 \text{ M} - 0.09992 \text{ M} = 8 \times 10^{-5} \text{ M}$$

$$[\text{H}_2]_{\text{eq}} = 0.150 - 3x = 0.150 - 3(8 \times 10^{-5} \text{ M}) = 0.1498 \text{ M}$$

$$[\text{NH}_3]_{\text{eq}} = 2x = 2(8 \times 10^{-5} \text{ M}) = 1.6 \times 10^{-4} \text{ M}$$

$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(1.6 \times 10^{-4})^2}{(0.09992)(0.1498)^3} = 7.62 \times 10^{-5}$$