Worksheet 18 - Equilibrium

Balance the following reaction, and use it to answer the following 8 questions:

$$\_\_\_ \; N_2\,(g) + \_\_\_ \; H_2\,(g) \leftrightarrows \_\_\_ \; NH_3\,(g)$$

1. Starting with **0.500 M N_2** and **0.800 M H_2**, the reaction is allowed to proceed until it reaches equilibrium. In the absence of products, the value of Q, the reaction quotient, is zero, less than any possible value of *K*.

   The **$[NH_3]_{eq}$** = **0.150 M** at 250°C. Determine ***K*** for this reaction at this temperature.

   a) Start by filling in the ICE table with the **Initial** concentrations.

	$[N_2]$ (M)	$[H_2]$ (M)	$[NH_3]$ (M)
Initial	_____	_____	_____
Change	_____	_____	_____
Equilibrium	_____	_____	_____

   b) Next, fill in the **Changes** in the system, in terms of **x**, which will determined by the **stoichiometry**. Reactants are used up (-) and products are formed (+).

   c) Next, write **equilibrium** expressions, in terms of initial concentrations and the changes that occur (x).

   d) Solve for **x** from the information you have.

   x = ______

   e) Write the **equation** for ***K***, the equilibrium constant, in terms of the **concentrations of reactants** and **products.**

   K =

   f) Finally, determine the **numerical value** of the equilibrium constant, ***K*** at this temperature.

   K =

2. Convert $\boldsymbol{K}$, based on concentration, to $\boldsymbol{K_P}$, based on partial pressures. The term $\boldsymbol{\Delta n}$ means the number of moles of gas in the products minus the number of moles of gas in the reactants. If $\Delta n = 0$, $K = K_P$.

$\boldsymbol{\Delta n}$ =

$\boldsymbol{K_P = K(RT)^{\Delta n}}$ =

3. Calculate the **partial pressures** of N_2, H_2 and NH_3 assuming a **10.0 L** volume, at 250°C.

a) First calculate the moles of each gas present at equilibrium.

mol N_2 =

mol H_2 =

mol NH_3 =

b) Now find the **total pressure**, using the ideal gas law.

P_{tot} = ________

c) Find the partial pressures of each of the gases using their mol fractions:

P_{N2} = (mol N_2/mol total) x P_{tot} = _____

P_{H2} = (mol H_2/mol total) x P_{tot} = _____

P_{NH3} = (mol NH_3/mol total) x P_{tot} = _____

4. Calculate $\boldsymbol{K_P}$ using these values. Is it the same as your previous answer, calculated from $\boldsymbol{K}$ in question 2?

$$K_P = \frac{P_{prod}^{\,p}}{P_{react}^{\,r}} =$$

5. Now, imagine that the volume of the container increases from **10.0 L** to **50.0 L**. Initially, there will be the same number of **moles** of each of the gases that were present at equilibrium. However, they are now in a larger volume, so their concentrations will change.

a) Determine the new concentrations of N_2, H_2 and NH_3.

$[N_2]$ = ______

$[H_2]$ = ______

$[NH_3]$ = ______

b) Using these new **initial concentrations**, determine the value of Q, the reaction quotient

$$Q = \frac{[\text{product}]_i^p}{[\text{reactant}]_i^r} =$$

c) Compare Q to *K*. Which way will the reaction proceed, under these new experimental conditions?

d) Compare the number of moles of gas on the reactant and product sides of the equation.

$n_{reactant}$ = ______ $n_{product}$ = ______

Increasing the volume of the system shifts the equilibrium toward the **reactants** / **products**, which has a **larger** / **smaller** number of moles of gas.

6. Start the reaction with **1.00 M N_2**, **1.00 M H_2** and **1.00 M NH_3** at the same temperature, 250^o C What is the value of **Q** under these initial conditions?

Q =

Compare Q and *K*. Which way will this reaction proceed?

Which compound will **decrease** in concentration?

Complete the ICE table for this set of initial conditions.

	$[N_2]$ (M)	$[H_2]$ (M)	$[NH_3]$ (M)
Initial	_1.00_	_1.00_	_1.00_
Change	______	______	______
Equilibrium	______	______	______

Write the expression for *K*.

K =

7. The equilibrium constant of this reaction changes with temperature. At T = 500^{o}C, $\boldsymbol{K_P = 3.6 \times 10^{-2}}$ and at 400^{o}C, $\boldsymbol{K_P = 41}$.

a) Raising the temperature **increases** or **decreases** K_p.

b) Raising the temperature drives the _______ reaction.

c) Should heat be treated as a **product** or a **reactant**?

d) Is this an **endothermic** or **exothermic** reaction?

8. In the original set of calculations, the equilibrium concentrations of the reactants and product were:

$[NH_3]_e = 0.150$ M $[N_2]_e = 0.425$ M $[H_2]_e = 0.575$ M

and $K_P = 1.51 \times 10^{-4}$ at 250^o C.

Assume a 1.00 L container. What is the effect of adding **1.00 moles** of **argon** to this container? It will not participate in the reaction, but it will change the partial pressures of all of the species.

Calculate the total number of moles of gas in the container.

$n_{total} =$

Calculate the mol fraction, X, of each of the species after the addition of the argon.

$X_{NH3} =$ _______

$X_{N2} =$ _______

$X_{H2} =$ _______

$X_{Ar} =$ _______

Calculate the pressure of this system, at 250^o C.

Calculate the partial pressures of each of the species.

$P_{NH3} =$ _______

$P_{N2} =$ _______

$P_{H2} =$ _______

Calculate K_P using these new equilibrium partial pressures.
What is the effect on the equilibrium of adding an inert gas?