AP CHEMISTRY

TOPIC 6: EQUILIBRIUM, REVIEW

Day 71:

- Chemical Equilibrium
- Heterogeneous Equilibria
- Applications of the Equilibrium Constant
- Solving Equilibrium Problems
- Equilibrium position
- Reaction Quotient
- Calculating Equilibrium Pressures
- Le Chatelier's Principle.
- Equilibrium expression
- Equilibrium (pressures)

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1. For the following process at 700. ^{0}C, what is the partial pressure of the gases at equilibrium if the total pressure is 0.750 atm? Carbon dioxide has a partial pressure of 0.201 atm.

$$C_{(s)} + CO_{2\,(g)} \leftrightarrow 2\,CO_{(g)} \qquad K_p = 1.50 \text{ atm}$$

Answers

$$K_p = \frac{P_{CO}^{\ 2}}{P_{CO_2}} = 1.50 \ ; \ \frac{P_{CO}^{\ 2}}{(0.201)} = 1.50 \ ; \ P_{CO}^{\ 2} = (1.50\ atm)(0.201\ atm) = 0.3015\ atm^2$$

$$P_{CO} = \sqrt{0.3015\ atm^2} = 0.549\ atm$$

2. Calculate the equilibrium constant, *K*, for the following reaction at 25.0 ^{0}C if the equilibrium concentrations are [Cl_2] = 0.371 *M*, [F_2] = 0.194 *M*, and [ClF] = 1.02 *M*.

$$Cl_{2\,(g)} + F_{2\,(g)} \leftrightarrow 2\,ClF_{(g)}$$

Answers

$$K_p = \frac{[\ ClF\ ]^2}{[\ Cl_2\ ][\ F_2\ ]} = \frac{(\ 1.02\ M\ )^2}{(\ 0.371\ M\ )(\ 0.194\ M\ )} = 14.5$$

3. Hypobromous acid, HOBr, dissociates in water according to the following reaction:

$$HOBr_{(aq)} \leftrightarrow OBr^{-1}_{(aq)} + H^{+1}_{(aq)} \qquad K = 2.06 \times 10^{-9} \text{ at } 25.0\ ^0C$$

Calculate the [H^{+1}] of a solution originally 1.25 *M* in HOBr.

Answers:

	[HOBr]	↔	[OBr^{-1}]	+	[H^{+1}]
I	1.25 mol / L		0		0
C	- *x*		+ *x*		+ *x*
E	1.25 - *x*		*x*		*x*

$$2.06 \times 10^{-9} = \frac{[\ OBr^{-1}\ ][\ H^{+1}\ ]}{[\ HOBr\ ]} = \frac{x^2}{(1.25 - x)} = \frac{x^2}{1.25} \quad (\text{assuming } 1.25 - x \approx 1.25)$$

$$(\ 1.25\ )(\ 2.06 \times 10^{-9}\ ) = x^2 \ ; \ \sqrt{2.575 \times 10^{-9}} = x = 5.07 \times 10^{-5}$$

$[\ H^{+1}\ ] = 5.07 \times 10^{-5} = x$, Assumption is great !!!

Other concentrations: $[\ H^{+1}\ ] = [\ OBr^{-1}\ ] = x = 5.07 \times 10^{-5}\ M \quad [\ HOBr\ ] = 1.25\ M$

4. The reaction of methane with water is given by the following equation:

$$CH_{4\,(g)} + H_2O_{(l)} \leftrightarrow CO_{(l)} + 3\,H_{2\,(g)} \qquad K = 5.67 \qquad \Delta H^0 = -350 \text{ kJ}$$

Predict the direction that the system will shift in order to reach equilibrium given the following situations.

Answers:

a.	$Q = 11.85$	$Q > K$, Shift to left (toward reactants)
b.	$Q = 3.8 \times 10^{-4}$	$Q < K$, Shift to right (toward products)
c.	*water is added*	No shift, water is a pure liquid
d.	*methane is reduced*	Shift to left (toward reactants)
e.	*energy is added*	Shift to left (toward reactants)
f.	*container's volume is reduced*	Shift to left (toward reactants)

5. The equilibrium constant is 9.30 atm $^{-2}$ at 25.0 ^{0}C for the reaction:

$$N_{2\,(g)} + 6\,HCl_{(g)} \leftrightarrow 2\,NH_{3\,(g)} + 3\,Cl_{2\,(g)}$$

The partial pressures for the gases are: P_{N_2} = 2.58 atm, P_{HCl} = 0.555 atm, P_{NH_3} = 1.45 atm. P_{Cl_2} = 0.750 atm,.
For this set of conditions, is the system at equilibrium (Show all work)? If not at equilibrium, in which direction will the system shift?

Answers:

$$Q = \frac{\left(P_{NH_3}{}^2\right)\left(P_{Cl_2}{}^3\right)}{\left(P_{N_2}\right)\left(P_{HCl}{}^6\right)} = \frac{(1.45\ atm)^2\,(0.750\ atm)^3}{(2.58\ atm)\,(0.555\ atm)^6} = 11.8\ atm^{-2}$$

$Q > K$
Shift to the Left

6. At 25 ^{0}C, K_p = 9.30 atm $^{-2}$ for the reaction (same as in question # 5):

$$N_{2\,(g)} + 6\,HCl_{(g)} \leftrightarrow 2\,NH_{3\,(g)} + 3\,Cl_{2\,(g)}$$

what is the value for K_c at this temperature.

Answers:

$$\Delta n = (2+5)-(1+6) = -2$$

$$K_p = K_c(RT)^{\Delta n} \quad \text{re-write as:} \quad K_c = \frac{K_p}{(RT)^{-2}} = K_p(RT)^{2}$$

$$K_C = K_p(RT)^{2} = \left(\frac{9.30}{atm^2}\right)\left(\left(\frac{0.0821\ atm\ L}{mol\ K}\right)(298\ K)\right)^2$$

$$K_C = \left(\frac{9.30}{atm^2}\right)\left(\left(\frac{0.00674\ atm^2\ L^2}{mol^2\ K^2}\right)(88804\ K^2)\right)$$

$$K_C = 5.57\times10^3\ \frac{L^2}{mol^2}$$