

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)

1. For the following process at 700. °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.

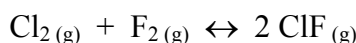


Answers

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

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

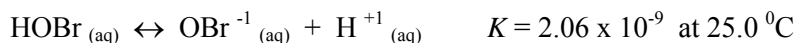
2. Calculate the equilibrium constant, K , for the following reaction at 25.0 °C if the equilibrium concentrations are $[Cl_2] = 0.371 M$, $[F_2] = 0.194 M$, and $[ClF] = 1.02 M$.



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:



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

Answers:

	[HOBr]	\leftrightarrow	[OBr ⁻¹]	+	[H ⁺¹]
I	1.25 mol / L		0		0
C	- x		+ x		+ x
E	1.25 - x		x		x

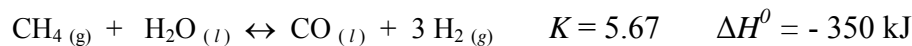
$$2.06 \times 10^{-9} = \frac{[OBr^{-1}][H^{+1}]}{[HOBr]} = \frac{x^2}{(1.25 - x)} = \frac{x^2}{1.25} \quad (\text{assuming } 1.25 - x \approx 1.25)$$

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

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

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

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

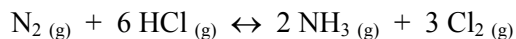


Predict the direction that the system will shift in order to reach equilibrium given the following situations.

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. <i>water is added</i>	No shift, water is a pure liquid
d. <i>methane is reduced</i>	Shift to left (toward reactants)
e. <i>energy is added</i>	Shift to left (toward reactants)
f. <i>container's volume is reduced</i>	Shift to left (toward reactants)

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



The partial pressures for the gases are: $P_{\text{N}_2} = 2.58 \text{ atm}$, $P_{\text{HCl}} = 0.555 \text{ atm}$, $P_{\text{NH}_3} = 1.45 \text{ atm}$, $P_{\text{Cl}_2} = 0.750 \text{ atm}$.

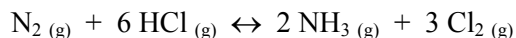
For this set of conditions, is the system at equilibrium (Show all work)? If not at equilibrium, in which direction will the system shift?

Answers:

$$Q = \frac{(P_{\text{NH}_3})^2 (P_{\text{Cl}_2})^3}{(P_{\text{N}_2}) (P_{\text{HCl}})^6} = \frac{(1.45 \text{ atm})^2 (0.750 \text{ atm})^3}{(2.58 \text{ atm}) (0.555 \text{ atm})^6} = 11.8 \text{ atm}^{-2}$$

$Q > K$
Shift to the Left

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



what is the value for K_c at this temperature.

Answers:

$$\Delta n = (2+3) - (1+6) = -2$$

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

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

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

$$K_c = 5.57 \times 10^3 \frac{\text{L}^2}{\text{mol}^2}$$