

Unit 8: Physical and Chemical Equilibrium

26. If there is no change in concentration, why is the equilibrium state considered dynamic ?

Rate of forward reaction = Rate of backward reaction

(Chemical reactions which are reversible do not cease, when equilibrium is attained. At equilibrium the forward and the backward reactions are proceeding at the same rate and no macroscopic change is observed. So chemical equilibrium is in a state of dynamic equilibrium.)

27. For a given reaction at a particular temperature, the equilibrium constant has constant value. Is the value of Q also constant ? Explain.

K_c and Q are constant at equilibrium both are temperature dependent. When K_c is constant at given temperature, Q also constant. (No. The value of Q is determined by the concentration of product and reactants that are not necessarily equilibrium concentration. Thus its value is not a constant. At equilibrium $K_c = Q$)

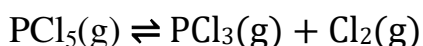
28. What the relation between K_p and K_c . Give one example for which K_p is equal to K_c .

Relation between K_p and K_c $K_p = K_c (RT)^{\Delta n_g}$

When $\Delta n_g = 0$; $K_p = K_c (RT)^0 = K_c$

Example : $H_2(g) + I_2 \rightleftharpoons 2HI(g)$

29. For a gaseous homogeneous reaction at equilibrium, number of moles of products are greater than the number of moles of reactants. Is K_c is larger or smaller than K_p .



$\Delta n_g = \text{No. of moles of product} - \text{No. of moles of reactant}$

$$\Delta n_g = 2 - 1 = 1$$

When $\Delta n_g = +ve$

$$K_p = K_c (RT)^{+ve} ; K_p > K_c$$

30. When the numerical value of the reaction quotient (Q) is greater than the equilibrium constant (K), in which direction does the reaction proceed to reach equilibrium ?

If $Q > K_c$, the reaction will proceed in the **reverse direction** i.e., formation of reactants.

31. For the reaction, $A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$; ΔH is $-ve$. the following molecular scenes represent different reaction mixture (A – green, B – blue)

- Calculate the equilibrium constant K_p and (K_c).
- For the reaction mixture represented by scene (x), (y) the reaction proceed in which directions ?
- What is the effect of increase in pressure for the mixture at equilibrium.

$$K_c = \frac{[AB]^2}{[A_2][B_2]} \quad A - \text{green} ; B - \text{blue}$$

Given that 'V' is constant (closed system) At equilibrium,

$$K_c = \frac{\left[\frac{4}{V}\right]^2}{\left(\frac{2}{V}\right)\left(\frac{2}{V}\right)} = \frac{16}{4} = 4$$

$$K_p = K_c (RT)^{\Delta n} ; K_p = 4(RT)^0 = 4$$

At stage 'x'

$$Q = \frac{\left(\frac{6}{V}\right)^2}{\left(\frac{2}{V}\right)\left(\frac{1}{V}\right)} = \frac{36}{2} = 18 ; Q > K_c \text{ (ie.),}$$

reverse reaction is favoured At Stage 'y'

$$Q = \frac{\left(\frac{3}{V}\right)^2}{\left(\frac{3}{V}\right)\left(\frac{3}{V}\right)} = \frac{9}{3 \times 3} = 1 ; K_c > Q \text{ (ie.),}$$

forward reaction is favoured

32. State Le-Chatelier principle.

It states that " if a system at equilibrium is disturbed, then the system shifts itself in a direction that nullifies the effect of that disturbance."

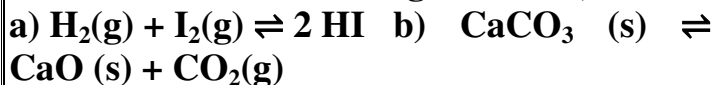
Dedication!

Determination!!

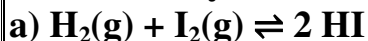
Distinction!!!

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33. Consider the following reactions,



c) $\text{S}(\text{s}) + 3\text{F}_2 (\text{g}) \rightleftharpoons \text{SF}_6 (\text{g})$ In each of the above reaction find out whether you have to increase (or) decrease the volume to increase the yield of the product.

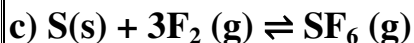


According to Le Chateliers principle increase in pressure will shift the equilibrium a direction that has lesser number of moles.

In the above equilibrium, pressure and volume has no effect, since no. of product is equal to no. of moles of reactant.



In this equilibrium no. of moles of product is greater than no. of moles of reactant. So increase of pressure will shift the equilibrium towards the backward reaction. To increase the yield of product, pressure should be decrease and hence volume increases.



In the above equilibrium increase in pressure favours formation of product hence volume should be decreased.

34. State law of mass action.

The law states that, “At any instant, the rate of a chemical reaction at a given temperature is directly proportional to the product of the active masses of the reactants at that instant”.

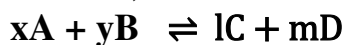
Rate \propto [Reactant]^x

Where, x is the stoichiometric coefficient of the reactant.

35. Explain how will you predict the direction of a equilibrium reaction.

From the knowledge of equilibrium constant, it is possible to predict the direction in which the net reaction is taking place for a given concentration or partial pressure of reactants and products.

Consider a general homogeneous reversible reaction,

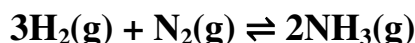


For the above reaction under nonequilibrium conditions, reaction quotient ‘Q’ is defined as the ratio of the product of active masses of reaction products raised to the respective stoichiometric coefficients in the balanced chemical equation to that of the reactants.

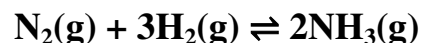
Under non-equilibrium conditions, the reaction quotient Q can be calculated using the following expression. As the reaction proceeds, there is a continuous change in the concentration of reactants and products and also the Q value until the reaction reaches the equilibrium. At equilibrium Q is equal to K_c at a particular temperature. Once the equilibrium is attained, there is no change in the Q value. By knowing the Q value, we can predict the direction of the reaction by comparing it with K_c.

- If $Q = K_c$, the reaction is in equilibrium state.
- If $Q > K_c$, the reaction will proceed in the reverse direction i.e., formation of reactants.
- If $Q < K_c$, the reaction will proceed in the forward direction i.e., formation of products.

36. Derive a general expression for the equilibrium constant K_p and K_c for the reaction



Let us consider the formation of ammonia in which, ‘a’ moles nitrogen and ‘b’ moles hydrogen gas are allowed to react in a container of volume V. Let ‘x’ moles of nitrogen react with 3x moles of hydrogen to give 2x moles of ammonia.



| | N_2 | H_2 | NH_3 |
|-------------------------|--------------|--------------|---------------|
| Initial number of moles | a | b | 0 |

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| | | | |
|---|-----------------|------------------|----------------|
| number of moles reacted | x | 3x | 0 |
| Number of moles at equilibrium | a-x | b-3x | 2x |
| Active mass or molar concentration at equilibrium | $\frac{a-x}{V}$ | $\frac{b-3x}{V}$ | $\frac{2x}{V}$ |

Applying law of mass action,

$$K_C = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_C = \frac{\left(\frac{2x}{V}\right)^2}{\left(\frac{a-x}{V}\right)\left(\frac{b-3x}{V}\right)^3} = \frac{\left(\frac{4x^2}{V}\right)}{\left(\frac{a-x}{V}\right)\left(\frac{b-3x}{V}\right)^3} = \frac{4x^2V^2}{(a-x)(b-3x)^2}$$

The equilibrium constant K_P can also be calculated as follows:

$$K_P = K_C (RT)^{(\Delta n_g)}$$

$$\Delta n_g = n_p - n_r = 2 - 4 = -2$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^2} (RT)^{-2}$$

Total number of moles at equilibrium,

$$n = a-x + b-3x + 2x = a+b-2x$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^2} \left(\frac{PV}{n}\right)^{-2}$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^2} \left(\frac{n}{PV}\right)^2$$

$$K_P = \frac{4x^2V^2}{(a-x)(b-3x)^2} \left(\frac{a+b-2x}{PV}\right)^2$$

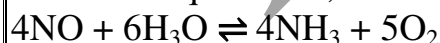
$$K_P = \frac{4x^2V(a+b-2x)^2}{P^2(a-x)(b-3x)^3}$$

37. Write a balanced chemical equation for a equilibrium reaction for which the equilibrium constant is given by expression

$$K_C = \frac{[NH_3]^4[O_2]^5}{[NO]^4[H_2O]^6}$$

$$K_C = \frac{[NH_3]^4[O_2]^5}{[NO]^4[H_2O]^6}$$

Chemical equation is,



38. What is the effect of added inert gas on the reaction at equilibrium.

When an inert gas (i.e, a gas which does not react with any other species involved in equilibrium) is added to an equilibrium system

at constant volume, the total number of moles of gases present in the container increases, that is, the total pressure of gases increases. The partial pressure of the reactants and the products or the molar concentration of the substance involved in the reaction remains unchanged. Hence at constant volume, addition of inert gas has no effect on equilibrium.

39. Derive the relation between K_P and K_C .

Let us consider the general reaction in which all reactants and products are ideal gases.



The equilibrium constant, K_C is, $K_C = \frac{[C]^l[D]^m}{[A]^x[B]^y}$

---- (1)

$$\text{And } K_P \text{ is, } K_P = \frac{P_C^l P_D^m}{P_A^x P_B^y} \text{ ----}$$

(2)

The ideal gas equation is

$$PV = nRT \text{ or } P = \frac{n}{V} RT$$

Since Active mass = molar concentration = n/V

P = active mass $\times RT$

Based on the above expression the partial pressure of the reactants and products can be expressed as,

$$P_A^x = [A]^x [RT]^x$$

$$P_B^y = [B]^y [RT]^y$$

$$P_C^l = [C]^l [RT]^l$$

$$P_D^m = [D]^m [RT]^m$$

On substitution in Eqn. 2,

$$K_P = \frac{[C]^l [RT]^l [D]^m [RT]^m}{[A]^x [RT]^x [B]^y [RT]^y} \text{ ---- (3)}$$

$$K_P = \frac{[C]^l [D]^m [RT]^{l+m}}{[A]^x [B]^y [RT]^{x+y}}$$

$$K_P = \frac{[C]^l [D]^m}{[A]^x [B]^y} RT^{(l+m)-(x+y)} \text{ ---- (4)}$$

By comparing equation (1) and (4), we get

$$K_P = K_C (RT)^{(\Delta n_g)} \text{ ---- (5)}$$

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where, Δn_g is the difference between the sum of number of moles of products and the sum of number of moles of reactants in the gas phase.

40. One mole of PCl_5 is heated in one litre closed container. If 0.6 mole of chlorine is found at equilibrium, calculate the value of equilibrium constant.

$$\begin{aligned}\text{Given that } [PCl_5]_{\text{initial}} &= \frac{1 \text{ mole}}{1 \text{ dm}^3} \\ [Cl_2]_{\text{eq}} &= 0.6 \text{ mole dm}^{-3} \\ PCl_5 &\rightleftharpoons PCl_3 + Cl_2 \\ [PCl_3]_{\text{eq}} &= 0.6 \text{ mole dm}^{-3} \\ [PCl_5]_{\text{eq}} &= 0.4 \text{ mole dm}^{-3} \\ K_C &= \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{0.6 \times 0.6}{0.4} = \mathbf{0.9}\end{aligned}$$

41. For the reaction $SrCO_3(s) \rightleftharpoons SrO(s) + CO_2(g)$, the value of equilibrium constant $K_p = 2.2 \times 10^{-4}$ at 1002 K. Calculate K_C for the reaction.

For the reaction, $SrCO_3(s) \rightleftharpoons SrO(s) + CO_2(g)$

$$\begin{aligned}\Delta n_g &= 1 - 0 = 1 \\ \therefore K_p &= K_C (RT) \\ 2.2 \times 10^{-4} &= K_C (0.0821) (1002) \\ K_C &= 2.674 \times 10^{-6}\end{aligned}$$

42. To study the decomposition of hydrogen iodide, a student fills an evacuated 3 litre flask with 0.3 mol of HI gas and allows the reaction to proceed at 500 °C. At equilibrium he found the concentration of HI which is equal to 0.05 M. Calculate K_C and K_p .

$$\begin{aligned}V &= 3L \quad ; \quad [HI]_{\text{initial}} = \frac{0.3 \text{ mol}}{3L} = 0.1M \\ [HI]_{\text{eq}} &= 0.05M \\ 2HI(g) &\rightleftharpoons H_2(g) + I_2(g)\end{aligned}$$

| | HI(g) | H ₂ (g) | I ₂ (g) |
|---------------------------|-------|--------------------|--------------------|
| Initial Concentration | 0.1 | - | - |
| Reacted | 0.05 | - | - |
| Equilibrium Concentration | 0.05 | 0.025 | 0.025 |

$$K_C = \frac{[H_2][I_2]}{[HI]^2} = K_C = \frac{0.025 \times 0.025}{0.05 \times 0.05} =$$

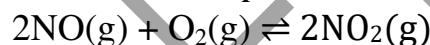
0.25

$$K_P = K_C (RT)^{\Delta n_g}$$

$$\Delta n_g = 2 - 2 = 0$$

$$K_P = 0.25 (RT)^0 = \mathbf{0.25}$$

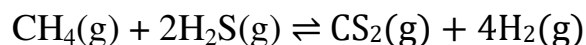
43. Oxidation of nitrogen monoxide was studied at 200 °C with initial pressures of 1 atm NO and 1 atm of O₂. At equilibrium partial pressure of oxygen is found to be 0.52 atm calculate K_p value.



| | NO | O ₂ | NO ₂ |
|------------------------------|------|----------------|-----------------|
| Initial Pressure | 1 | 1 | - |
| Reacted | 0.96 | 0.48 | - |
| Equilibrium partial pressure | 0.04 | 0.52 | 0.96 |

$$K_P = \frac{(P_{NO_2})^2}{(P_{NO})^2(P_{O_2})} = \frac{0.96 \times 0.96}{0.04 \times 0.04 \times 0.52} = \mathbf{1.017} \times 10^3$$

44. 1 mol of CH₄, 1 mole of CS₂ and 2 mol of H₂S are 2 mol of H₂ are mixed in a 500 ml flask. The equilibrium constant for the reaction $K_C = 4 \times 10^{-2} \text{ mol}^2 \text{ lit}^{-2}$. In which direction will the reaction proceed to reach equilibrium ?



$$K_C = 4 \times 10^{-2} \text{ mol lit}^{-2}$$

$$\text{Volume} = 500\text{ml} = \frac{1}{2} \text{ L}$$

$$\begin{aligned}[CH_4]_{\text{in}} &= \frac{1 \text{ mol}}{\frac{1}{2} \text{ L}} = 2 \text{ mol L}^{-1} & [CS_2]_{\text{in}} &= \frac{1 \text{ mol}}{\frac{1}{2} \text{ L}} = 2 \text{ mol L}^{-1} \\ [H_2S]_{\text{in}} &= \frac{2 \text{ mol}}{\frac{1}{2} \text{ L}} = 4 \text{ mol L}^{-1} & [H_2] &= \frac{2 \text{ mol}}{\frac{1}{2} \text{ L}} = 4 \text{ mol L}^{-1}\end{aligned}$$

$$Q = \frac{[CS_2][H_2]^4}{[CH_4][H_2S]^2} = \frac{2 \times (4)^4}{(2)(2)^2} = 64; \quad Q > K_C$$

45. At particular temperature $K_C = 4 \times 10^{-2}$ for the reaction

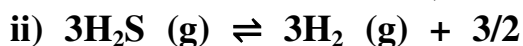
$H_2S(g) \rightleftharpoons H_2(g) + \frac{1}{2} S_2(g)$ Calculate KC for each of the following reaction i) $2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$

Dedication!

Determination!!

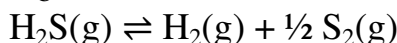
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$\text{S}_2(\text{g})$

$K_C = 4 \times 10^{-2}$ for the reaction,



$$K_C = \frac{[\text{H}_2][\text{S}_2]^{\frac{1}{2}}}{[\text{H}_2\text{S}]}$$

$$4 \times 10^{-2} = \frac{[\text{H}_2][\text{S}_2]^{\frac{1}{2}}}{[\text{H}_2\text{S}]}$$

For the reaction, i) $2\text{H}_2\text{S}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{S}_2(\text{g})$

$$K_C = \frac{[\text{H}_2]^2[\text{S}_2]}{[\text{H}_2\text{S}]^2} = (4 \times 10^{-2})^2 = 16 \times 10^{-4}$$

For the reaction, ii) $3\text{H}_2\text{S}(\text{g}) \rightleftharpoons 3\text{H}_2(\text{g}) + \frac{3}{2}\text{S}_2(\text{g})$

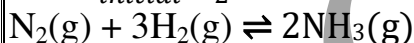
$$K_C = \frac{[\text{H}_2]^3[\text{S}_2]^{\frac{3}{2}}}{[\text{H}_2\text{S}]^3} = (4 \times 10^{-2})^2 = 64 \times 10^{-6}$$

46. 28 g of Nitrogen and 6 g of hydrogen were mixed in a 1 litre closed container. At equilibrium 17 g NH_3 was produced. Calculate the weight of nitrogen, hydrogen at equilibrium.

Given $m_{\text{N}_2} = 28\text{g}$; $m_{\text{H}_2} = 6\text{g}$; $V = 1\text{L}$

$$(n_{\text{N}_2})_{\text{initial}} = \frac{28}{28} = 1 \text{ mol}$$

$$(n_{\text{H}_2})_{\text{initial}} = \frac{6}{2} = 3 \text{ mol}$$



| | $\text{N}_2(\text{g})$ | $\text{H}_2(\text{g})$ | $\text{NH}_3(\text{g})$ |
|---------------------------|------------------------|------------------------|-------------------------|
| Initial Concentration | 1 | 3 | - |
| Reacted | 0.5 | 1.5 | - |
| Equilibrium Concentration | 0.5 | 1.5 | 1 |

$$[\text{NH}_3] = \left(\frac{17}{17}\right) = 1 \text{ mol}$$

Weight of N_2 = (no. of moles of N_2) \times molar mass of N_2

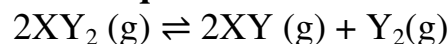
$$= 0.5 \times 28 = 14\text{g}$$

Weight of H_2 = (no. of moles of H_2) \times molar mass of H_2

$$= 1.5 \times 2 = 3\text{g}$$

47. The equilibrium for the dissociation of XY_2 is given as, $2\text{XY}_2(\text{g}) \rightleftharpoons 2\text{XY}(\text{g}) + \text{Y}_2(\text{g})$

if the degree of dissociation x is so small compared to one. Show that $2 K_P = P X^3$ where P is the total pressure and K_P is the dissociation equilibrium constant of XY_2 .



| | XY_2 | XY | Y_2 |
|-----------------------------|-------------------|-------------|--------------|
| Initial no. of moles | 1 | - | - |
| No. of moles dissociated | X | - | - |
| No. of moles at equilibrium | $(1 - X) \cong 1$ | X | X/2 |

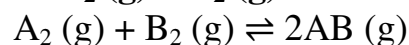
Total no. of moles = $1 - X + X + X/2 = 1 + X/2 \cong 1$

[therefore Given that $x \ll 1$; $1 - x \cong 1$ and $1 + X/2 \cong 1$]

$$K_P = \frac{(P_{\text{XY}})^2(P_{\text{Y}_2})}{(P_{\text{XY}_2})^2} = \frac{\left(\frac{x}{1} \times P\right)^2 \left(\frac{x}{2} \times P\right)}{\left(\frac{1}{1} \times P\right)^2}$$

$$K_P = \frac{x^2 P^2 \times P}{2 P^2} = 2 K_P = X^3 P$$

48. A sealed container was filled with 1 mol of $\text{A}_2(\text{g})$, 1 mol $\text{B}_2(\text{g})$ at 800 K and total pressure 1.00 bar. Calculate the amounts of the components in the mixture at equilibrium given that $K = 1$ for the reaction $\text{A}_2(\text{g}) + \text{B}_2(\text{g}) \rightleftharpoons 2\text{AB}(\text{g})$



| | A_2 | B_2 | AB |
|-----------------------------|--------------|--------------|-------------|
| Initial Concentration | 1 | 1 | - |
| No. Of moles dissociated | X | X | - |
| No. of moles at equilibrium | $1 - x$ | $1 - x$ | $2x$ |

Total no. of moles = $1 - x + 1 - x + 2x = 2$

$$K_P = \frac{(P_{\text{AB}})^2}{(P_{\text{A}_2})(P_{\text{B}_2})} = \frac{\left(\frac{2x}{2} \times P\right)^2}{\left(\frac{1-x}{2} \times P\right)\left(\frac{1-x}{2} \times P\right)}$$

$$K_P = \frac{4x^2}{(1-x)^2}$$

Given that $K_P = 1$; $\frac{4x^2}{(1-x)^2} = 1$

$$\Rightarrow 4x^2 = (1 - x)^2$$

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$$\Rightarrow 4x^2 = 1 + x^2 - 2x$$

$$3x^2 + 2x - 1 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-2 \pm \sqrt{4 - 4 \times 3 \times -1}}{2(3)} = \frac{-2 \pm \sqrt{4 + 12}}{6} = \frac{-2 \pm \sqrt{16}}{6} = \frac{-2 \pm 4}{6} = \frac{-2+4}{6}; \frac{-2-4}{6}$$

$$= \frac{2}{6}; \frac{-6}{6}$$

$x = 0.33; -1$ (not possible)

$$[A_2]_{eq} = 1 - x = 1 - 0.33 = 0.67$$

$$[B_2]_{eq} = 1 - x = 1 - 0.33 = 0.67$$

$$[AB]_{eq} = 2x = 2 \times 0.33 = 0.66$$

49. Deduce the Vant Hoff equation.

This equation gives the quantitative temperature dependence of equilibrium constant (K). The relation between standard free energy change (ΔG°) and equilibrium constant is

$$\Delta G^\circ = -RT \ln K \quad \text{----- (1)}$$

We know that

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \quad \text{----- (2)}$$

Substituting (1) in equation (2)

$$-RT \ln K = \Delta H^\circ - T\Delta S^\circ$$

Rearranging

$$\ln K = \frac{-\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R} \quad \text{----- (3)}$$

Differentiating equation (3) with respect to temperature,

$$\frac{d(\ln K)}{dT} = \frac{\Delta H^\circ}{RT^2} \quad \text{----- (4)}$$

Equation 4 is known as differential form of van't Hoff equation. On integrating the equation 4, between T_1 and T_2 with their respective equilibrium constants K_1 and K_2 .

$$\int_{K_1}^{K_2} d(\ln K) = \frac{\Delta H^\circ}{R} \int_{T_1}^{T_2} \frac{dT}{T^2}$$

$$[\ln K]_{K_1}^{K_2} = \frac{\Delta H^\circ}{R} \left[-\frac{1}{T} \right]_{T_1}^{T_2}$$

$$\ln K_2 - \ln K_1 = \frac{\Delta H^\circ}{R} - \left[\frac{1}{T_2} + \frac{1}{T_1} \right]$$

$$\ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left[\frac{T_2 - T_1}{T_2 T_1} \right]$$

$$\log \frac{K_2}{K_1} = \frac{\Delta H^\circ}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1} \right] \quad \text{----- (5)}$$

Equation 5 is known as integrated form of van't Hoff equation.

50. The equilibrium constant K_p for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 8.19×10^2 at 298 K and 4.6×10^{-1} at 498 K. Calculate ΔH° for the reaction.

$$K_{P_1} = 8.19 \times 10^2 \quad T_1 = 298K$$

$$K_{P_2} = 4.6 \times 10^{-2} \quad T_2 = 498K$$

$$\log \left(\frac{K_{P_2}}{K_{P_1}} \right) = \frac{\Delta H^\circ}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1} \right]$$

$$\log \left(\frac{4.6 \times 10^{-1}}{8.19 \times 10^2} \right) = \frac{\Delta H^\circ}{2.303 \times 8.314} \left(\frac{498 - 298}{498 \times 298} \right)$$

$$\frac{-3.2505 \times 2.303 \times 8.314 \times 498 \times 298}{200} = \Delta H^\circ$$

$$\Delta H^\circ = -46181 \text{ J mol}^{-1}$$

$$\Delta H^\circ = -46.18 \text{ kJ mol}^{-1}$$

51. The partial pressure of carbon dioxide in the reaction $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ is 1.017×10^{-3} atm at 500°C . Calculate K_p at 600°C for the reaction. ΔH for the reaction is 181 kJ mol^{-1} and does not change in the given range of temperature.

$$P_{CO_2} = 1.017 \times 10^{-3} \text{ atm}; T = 500^\circ \text{C}$$

$$K_p = P_{CO_2}$$

$$K_{P_1} = 1.017 \times 10^{-3}; T = 500 + 273 = 773K$$

$$K_{P_2} = ? \quad T = 600 + 273 = 873K$$

$$\Delta H^\circ = 181 \text{ kJ mol}^{-1}$$

$$\log \left(\frac{K_{P_2}}{K_{P_1}} \right) = \frac{\Delta H^\circ}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1} \right]$$

$$\log \left(\frac{K_{P_2}}{1.017 \times 10^{-3}} \right) = \frac{181 \times 10^3}{2.303 \times 8.314} \left(\frac{873 - 773}{873 \times 773} \right)$$

$$\log \left(\frac{K_{P_2}}{1.017 \times 10^{-3}} \right) = \frac{181 \times 10^3 \times 100}{2.303 \times 8.314 \times 873 \times 773}$$

$$\frac{K_{P_2}}{1.017 \times 10^{-3}} = \text{anti log of (1.40)}$$

$$\frac{K_{P_2}}{1.017 \times 10^{-3}} = 25.12$$

$$K_{P_2} = 25.12 \times 1.017 \times 10^{-3}$$

$$K_{P_2} = 25.54 \times 10^{-3}$$