Chapter 9: Chemical Equilibrium

Review Questions

1. Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression.

Solution

The equilibrium constant expression for pure liquids and solids is as follows:

$$[\text{Pure liquid}]\text{or}[\text{Pure solid}] = \frac{\text{Number of moles}}{\text{Volume}} = \frac{\text{Mass}}{\text{Molecular mass} \times \text{Volume}} = \frac{\text{Density}}{\text{Molecular mass}}$$

As the density of a pure liquid or pure solid is constant at constant temperature and molecular mass is also constant, therefore, their molar concentrations are constant and included into the equilibrium constant. Hence, the values of pure substances are not mentioned in the equilibrium constant expression.

2. What is the difference between a heterogeneous and homogeneous equilibrium?

Solution

In a homogeneous equilibrium, all of the reactants and products are in the same phase. In heterogeneous equilibrium, at least two different phases are found between the reactants and products.

3. The ester, ethyl acetate, is formed by the reaction of ethanol and acetic acid and the equilibrium is represented as

$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$

(a) Write the concentration ratio, Q_c, for this reaction. Note that water is not in excess and is not a solvent in this reaction.

(b) At 293 K, if one starts with 1.000 mol of acetic acid and 1.180 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.

(c) Starting with 0.500 mol of ethanol and 1.000 mol of acetic acid and maintaining it at 293 K, 0.214 mol of ethyl acetate is found after some time. Has equilibrium been reached?

(d) We do not use dilute aqueous solutions for this reaction. Why?

(e) Why is some concentrated sulphuric acid usually added to reaction mixture in the laboratory preparation of ethyl acetate?

(f) As the heat of reaction is nearly zero for this reaction, how will the equilibrium constant depend upon the temperature?

Solution

(a) For the given reaction,

$$Q_c = \frac{[CH_3COOC_2H_5(l)][H_2O(l)]}{[CH_3COO(l)][C_2H_5OH(l)]}$$

(b)

	$CH_3COOH(l)$	$+ C_2H_5OH(l)$	$\rightleftharpoons CH_3COOC_2H_5$	$+ H_2O$
Initial	1.000 mol	0.180 mol	0	0
At equilibrium	(1 − 0.71) mol	(0.180 − 0.171) mol	0.171 mol	0.171 mol
	0.829 mol	0.009 mol	0.171 mol	0.171 mol

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So, $K = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]} = \frac{0.171\,\text{mol} \times 0.171\,\text{mol}}{0.829\,\text{mol} \times 0.009\,\text{mol}} = 3.92\,\text{mol}$

(c) Here

	$CH_3COOH(l)$	$+ C_2H_5OH(l)$	$\rightleftharpoons CH_3COOC_2H_5$	$+ H_2O$
Initial	1.500 mol	1.000 mol	0	0
At equilibrium	(0.500 – 0.214)	(1.000 – 0.214)	0.214	0.214 mol
	0.286 mol	0.786 mol	0.214 mol	0.214 mol

So, $Q_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]} = \frac{0.214\,\text{mol} \times 0.214\,\text{mol}}{0.286\,\text{mol} \times 0.786\,\text{mol}} = 0.20$

As the value of Q_c (= 0.20) is less than the value of K (= 3.92), hence the equilibrium is not reached. The reaction is on the left.

So, the reactants are still reaching to form more products.

(d) This reaction is a reversible reaction in which H_2O is one of the products. In dilute solution, the concentration of water is high. High concentration of water (H_2O) will shift the equilibrium toward left (reactant side). As a result, no or very little amount of ester will be formed if dilute aqueous solutions are used in this reaction.

(e) Concentrated sulphuric acid acts as a catalyst in this reaction. Also, concentrated sulphuric acid is a strong dehydrating agent. So, it removes water (H_2O) as soon as it is formed. As a result, the equilibrium shifts to the right, and more amount of ester is formed.

(f) Variation in equilibrium constant for this reaction would be small with a change of temperature.

4. Define physical equilibrium with examples.

Solution

The state of a process in which properties such as temperature, pressure, and concentration of the system do not show any change with the passage of time is known as equilibrium. In all processes that attain equilibrium, two opposing processes are involved. Equilibrium is attained when the rates of the two opposing processes become equal. If the opposing processes involve only physical changes, the equilibrium is called physical equilibrium. The most common examples of physical equilibrium are those that involve phase transformation. For example,

Solid $\rightleftharpoons$ liquid Liquid $\rightleftharpoons$ gas Solid $\rightleftharpoons$ gas

5. For the exothermic formation of sulphur trioxide from sulphur dioxide and oxygen in the gas phase:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

K_p = 40.5 atm^{-1} at 900 K and $\Delta_r H$ = –198 kJ mol^{-1}

(a) Write the expression for the equilibrium constant for the reaction.

(b) At room temperature ($\approx$ 300 K) will K_p be greater than, less than, or equal to K_p at 900 K.

(c) How will the equilibrium be affected if the volume of the vessel containing the three gases is reduced, keeping the temperature constant. What happens?

(d) What is the effect of adding 1 mol of He (g) to a flask containing SO_2, O_2, and SO_3 at equilibrium at constant temperature.

Solution

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(a) The equilibrium constant for this reaction is written in terms of the partial pressures of the reactants and products.

$$K_p = \frac{p^2_{SO_3(g)}}{p^2_{SO_2(g)} \times p_{O_2(g)}}$$

(b) This reaction is exothermic. So, its equilibrium constant should increase with the lowering of temperature. Therefore, the value of K_p at 300 K will be greater than the value at 900 K.

(c) When the volume of the vessel is reduced, the volume of the reaction mixture will decrease. As a result, pressure of the gaseous mixture will increase. According to the Le Chatelier's principle, the system will move in a direction to undo the effect of the pressure increase. The system therefore will move in a direction so as to decrease the number of moles of the gaseous substance in the system. The number of moles decreases in going from reactants to the product side. Therefore, a decrease in the volume of the reacting system will shift the equilibrium to the right. That is, more $SO_3(g)$ will be formed from the reactants.

(d) Addition of helium to the reaction mixture at equilibrium under constant volume has no effect on the equilibrium.

6. Why does not a catalyst affect the position of equilibrium in a chemical reaction?

Solution

A catalyst affects the rate of the reaction but does not affect the position of the equilibrium. The catalyst affects both the forward and the reverse reaction to the same extent, so the ratios of the forward and the reverse reactions for the catalyzed and uncatalyzed reactions do not change.

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7. Under what conditions will the following reactions go in the forward direction?

(a) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + 45$ kcal.

(b) $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) + 27.8$ kcal.

(c) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) - x$ kcal.

Solution

(a) Low T, high p, excess of O_2 and SO_2
(b) Low T, high p, excess of NO and O_2
(c) High T, low p, excess of $PC1_5$

8. Consider the following equilibrium.

$$2NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

If you were converting between K_p and K_c, what value of Δn_g would you use?

Solution

The value for Δn_g is the difference between the number of moles of reactant gas and the number of moles of product gas:

Δn_g = (moles of gaseous products) – (moles of gaseous reactants)
= 2 mol – 0 mol = 2 mol

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9. How will the value of K_p for the following reactions be affected by an increase in temperature?

(a) $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g);\ \Delta H^o = -18\,kJ$

(b) $N_2O(g) + NO_2(g) \rightleftharpoons 3NO(g);\ \Delta H^o = +155.7\,kJ$

(c) $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g);\ \Delta H^o = -77.07\,kJ$

Solution

(a) This is an exothermic reaction. By increasing the temperature, the equilibrium shifts to the left and more reactants are formed; therefore, the denominator becomes larger and the K_p decreases.

(b) This is an endothermic reaction. By increasing the temperature, the equilibrium is shifted to the right and more products are formed, and the numerator increases and K_p increases.

(c) This is the same as part (a): an exothermic reaction and the increase in temperature shift the reaction to the left and increase the amount of the reactants and K_p decreases.

10. Write the equilibrium law for each of the following gaseous reactions in terms of molar concentrations and partial pressures.

(a) $2PCl_3(g) + O_2(g) \rightleftharpoons 2POCl_3(g)$

(b) $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$

(c) $2N_2H_4(g) + 2O_2(g) \rightleftharpoons 2NO(g) + 2H_2O(g)$

(d) $N_2H_4(g) + 6H_2O_2(g) \rightleftharpoons 2NO_2(g) + 8H_2O(g)$

(e) $SOCl_2(g) + H_2O(g) \rightleftharpoons SO_2(g) + 2HCl(g)$

Solution

(a) $K_c = \dfrac{[POCl_3]^2}{[PCl_3]^2[O_2]},\ K_p = \dfrac{(p_{POCl_3})^2}{(p_{PCl_3})^2(p_{O_2})},$

(b) $K_c = \dfrac{[SO_2]^2[O_2]}{[SO_3]^2},\ K_p = \dfrac{(p_{SO_2})^2(p_{O_2})}{(p_{SO_3})^2},$

(c) $K_c = \dfrac{[NO]^2[H_2O]^2}{[N_2H_4][O_2]^2},\ K_p = \dfrac{(p_{NO})^2(p_{H_2O})^2}{(p_{N_2H_4})(p_{O_2})^2},$

(d) $K_c = \dfrac{[NO_2]^2[H_2O]^8}{[N_2H_4][H_2O_2]^6},\ K_p = \dfrac{(p_{NO_2})^2(p_{H_2O})^8}{(p_{N_2H_4})(p_{H_2O_2})^6},$

(e) $K_c = \dfrac{[SO_2]^2[HCl]^2}{[SOCl_2][H_2O]},\ K_p = \dfrac{(p_{SO_2})^2(p_{HCl})^2}{(P_{SOCl_2})(p_{H_2O})},$

11. How will the position of equilibrium in the following reaction:

$$\text{Heat} + CH_4(g) + 2H_2S(g) \rightleftharpoons CS_2(g) + 4H_2(g)$$

be affected by:

(a) Adding $CH_4(g)$?

(b) Adding $H_2(g)$?

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(c) Removing $CS_2(g)$? (d) Decreasing the volume of the container?
(e) Increasing the temperature?

Solution

(a) The system shifts to the right to consume some of the added methane.
(b) The system shifts to the left to consume some of the added hydrogen.
(c) The system shifts to the right to make some more carbon disulphide.
(d) The system shifts to the left to decrease the amount of gaseous moles.
(e) The system shifts to the right to absorb some of the added heat.

12. Two 1.00 L bulbs are filled with 0.500 atm of $F_2(g)$ and $PF_3(g)$, as illustrated in the Fig. 8.5. At a particular temperature, K_p = 4.0 for the reaction of the gases to form $PF_5(g)$:

$$F_2(g) + PF_3(g) \rightleftharpoons PF_5(g)$$

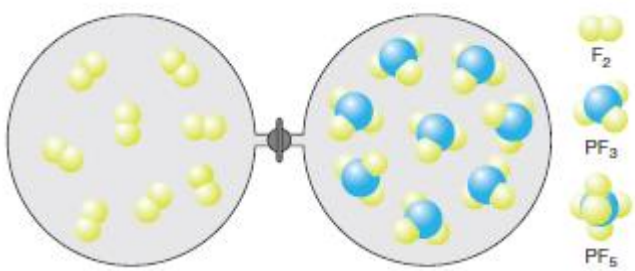
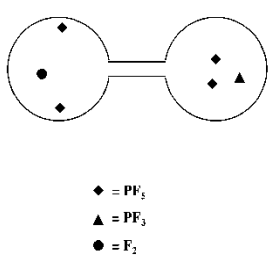


Figure 8.5

(a) The stopcock between the bulbs is opened and the pressure falls. Make a sketch of this apparatus showing the gas composition once the pressure is stable.
(b) List any chemical reactions that continue to occur once the pressure is stable.
(c) Suppose all the gas in the left bulb is forced into the bulb on the right and then the stopcock is closed. Make a second sketch to show the composition of the gas mixture after equilibrium is reached. Comment on how the composition of the one-bulb system differs from that for the two-bulb system.

Solution

(a)



(b) $F_2(g) + PF_3(g) \rightarrow PF_5(g)$
$PF_5(g) \rightarrow F_2(g) + PF_3(g)$

(c)

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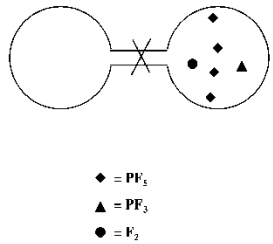

Because the volume is smaller, Le Chatelier's principle states that the equilibrium will favor the products. Consequently, equilibrium will be shifted to the right in a smaller volume. As can be seen here, a higher proportion of the reactants have reacted to form more PF_5 than the case in (a).

13. Use the following equilibria:

$2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g)$; $K_c = 9.5 \times 10^{-13}$

$CH_4(g) + H_2O(g) \rightleftharpoons CH_3OH(g) + H_2(g)$; $K_c = 2.8 \times 10^{-21}$

to calculate K_c for the reaction

$2CH_3OH(g) + H_2(g) \rightleftharpoons C_2H_6(g) + 2H_2O(g)$

Solution

If we reverse the second reaction and double it, we can then add the first and second reactions together to obtain the desired reaction. When we double the second reaction, we square the rate constant. When we reverse it, we invert the rate constant. Consequently, we get the following:

$$2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g); \quad K_c = 9.5 \times 10^{-13}$$

$$2CH_3OH(g) + 2H_2(g) \rightleftharpoons 2CH_4(g) + 2H_2O(g); K_c = 1.3 \times 10^{41}$$

Adding these equations, we get

$$2CH_3OH(g) + H_2(g) \rightleftharpoons C_2H_6(g) + 2H_2O(g); K_c = 1.2 \times 10^{29}$$

where K_c for the final reaction is the product of K_c for the first reaction and K_c for the "modified" second reaction.

14. Define the state of equilibrium. What are the general characteristics of the following equlibria?

(a) Solid–liquid (b) Liquid–vapor (c) Solid–vapor

Solution

Equilibrium represents the state of a process in which properties such as temperature, pressure, and concentration of the system do not show any change with the passage of time.

(a) Solid–Liquid equilibrium (melting of ice)

If some ice cubes along with some water at 0°C and normal atmospheric pressure are placed in a thermos flask so that no heat can enter or leave the system, the mass of ice and water is found to remain constant. However, the two opposing processes going on at equilibrium are melting of ice and freezing of water. At equilibrium

$$\text{Solid (ice)} \rightleftharpoons \text{Liquid (water)}$$

$$\text{Rate of melting of ice} = \text{Rate of freezing of water}$$

(b) Liquid–Vapor equilibrium (Evaporation of water in a closed vessel)

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Consider a closed vessel connected to a manometer having arrangement for evaluation and addition of liquid into it. Suppose the vessel is first evacuated. The level of mercury in both the limbs of the manometer will be the same. Now suppose water is added into the vessel and the whole apparatus is allowed to stay at room temperature. It is observed that the level of mercury in the left limb of the manometer begins to fall and that in the right limb begins to rise. After some time, the level becomes constant and the equilibrium is attained

$$H_2O(l) \rightleftharpoons H_2O(g)$$

$$\text{Rate of evaporation} = \text{Rate of condensation}$$

(c) Solid–vapor equilibrium

This type of equilibrium is attained for solids that undergo sublimation. For example, if solid iodine is placed in a closed vessel, violet vapors start appearing in the vessel whose intensity increases with time and ultimately it becomes constant. At this stage, equilibrium is attained, that is,

Rate of sublimation of solid iodine to vapor = Rate of condensation of iodine vapor to give solid iodine

The equilibrium may be represented as

$$I_2(s) \rightleftharpoons I_2(\text{vapor})$$

15. What are the general characteristics of equilibrium involving physical processes?

Solution

The general characteristics of equilibrium involving physical processes are as follows:

(a) Equilibria involving gases can be attained only in closed vessels.

(b) Equilibrium is dynamic in nature.

(c) Some observable properties of the system become constant at equilibrium.

(d) The concentration of the different substances becomes constant at constant temperature at equilibrium.

(e) The magnitude of the equilibrium constant represents the extent to which the process proceeds before equilibrium is attained.

16. What are reversible and irreversible reactions? Give one example of each.

Solution

A reaction in which not only the reactants react to form the products under certain conditions, but the products also react to form reactants under the same conditions is called a reversible reaction. In other words, a reaction that takes place not only in the forward direction but also in the backward direction under the same conditions is known as a reversible reaction. An example of reversible reaction is

$$3Fe(s) + 4H_2O(aq) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$$

If a reaction cannot take place in reverse direction, that is, the products formed do not react to give back the reactants under the same conditions, it is called an irreversible reaction. For example,

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_2(aq)$$

17. State the law of mass action.

Solution

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According to the law of mass action, the rate at which a substance reacts is proportional to its active mass, and hence, the rate of a chemical reaction is proportional to the product of the active masses of the reactants. The term active mass means the molar concentration, that is, number of moles dissolved per liter of the solution.

18. The equilibrium constant of a reaction is 2×10^{-3} at 25°C and 2×10^{-2} at 50°C. Is the reaction exothermic or endothermic?

Solution

As equilibrium constant has increased with temperature, the reaction is endothermic (proceeds in the forward direction).

19. Deduce the expression for equilibrium constant using the law of mass action.

Solution

This law states that: "At constant temperature the rate at which a substance reacts is proportional to its active mass, and the rate of a chemical reaction is proportional to the product of active masses of the reactants." The expression for active mass is

$$\text{Active mass (molar concentration)} = \frac{\text{Number of moles}}{\text{Volume in liters}}$$

Active mass of a gas or liquid means its molar concentration; active mass of a solid is always taken as unity. Law of mass action is applicable only to a reversible reaction. In other words, the law states that: "At constant temperature, the rate of a chemical reaction is directly proportional to the product of the molar concentrations of reacting species, with each concentration term raised to the power equal to the numerical coefficient of that species in the chemical equation." Let us consider a general reversible reaction:

$$aA + bB \rightleftharpoons cC + dD$$

Rate of forward reaction = $k_f\,[A]^a[B]^b$

Similarly, Rate of backward reaction = $k_b[C]^c[D]^d$

where k_f and k_b are the respective rate constants or velocity constant.

At equilibrium,

Rate of forward reaction = Rate of backward reaction

or $$k_f[A]^a[B]^b = k_b[C]^c[D]^d$$

therefore, $$\frac{k_f}{k_b} = \frac{[C]^c[D]^d}{[A]^a[B]^b} = K_c$$

where K_c is equilibrium constant. Conventionally, K_c is always equal to k_f/k_b.

20. What qualitative information can you obtain from the magnitude of equilibrium constant?

Solution

(a) Large value of equilibrium constant, K ($>10^3$) shows that forward reaction is favored, that is, concentration of products is much larger than that of the reactants at equilibrium.

(b) Intermediate value of K (10^{-3} to 10^3) shows that the concentration of the reactants and product is comparable.

(c) Low value of K ($<10^{-3}$) shows that backward reaction is favored, that is, concentration of reactants is much more than that of the products.

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21. For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, at 400 K, K_p = 41. Find the value of K_p for each of the following reactions at the same temperature:

(a) $2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$

(b) $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightleftharpoons NH_3(g)$

(c) $2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$

Solution

(a) It is the reverse of the given reaction. Hence, K_p = 1/41.
(b) It is obtained by dividing the given equation by 2. Hence, $K_p = (41)^{1/2}$
(c) It is obtained by multiplying the given equation by 2. Hence, $K_p = (41)^2$

22. Why does manufacture of ammonia by Haber's process require higher pressure, low temperature, and use of catalyst?

Solution

The reaction involving formation of ammonia is

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4 \text{ kJ mol}^{-1}$$
$$1\text{mol} \quad 3\text{mol} \quad 2\text{mol}$$

In this equilibrium reaction, the forward reaction is accompanied by a decrease in the total number of moles. If the pressure on the system is increased, then according to Le Chatelier's principle, the equilibrium will shift in the direction in where the pressure decreases. Hence, decrease in the number of moles takes place, which is in favor of formation of ammonia. Thus, higher the pressure, the better would be the yield of ammonia. Now, the forward reaction is exothermic. Therefore, according to Le Chatelier's principle, decrease in temperature will shift the reaction in the forward direction, that is, in favor of formation of ammonia. In this reaction, finely divided iron is used as a catalyst to achieve the equilibrium rapidly. Thus, manufacture of ammonia by Haber's process is favored by high pressure, low temperature, and use of catalyst.

23. Using Le Chatelier's principle, explain the effect of pressure on the solubility of gases in liquids.

Solution

Aerated water is the solution of CO_2 into water.

$$CO_2(s) + H_2O(l) \rightleftharpoons \text{Gas solution}$$

According to Le Chatelier's principle, high pressure shifts the equilibrium to the right, meaning reduced volume. Low pressure shifts the equilibrium to the left favoring the separation of the gas from the liquid as the volume increases. Therefore, CO_2 remains dissolved in water under pressure.

Numerical Problems

1. Nitrous oxide, N_2O, sometimes called laughing gas, has a strong tendency to decompose into nitrogen and oxygen following the equation:

$$2N_2O(g) \rightleftharpoons 2N_2(g) + O_2(g)$$

but the reaction is so slow that the gas appears to be stable at room temperature (25°C). The decomposition reaction has $K_c = 7.3 \times 10^{34}$. What is the value of K_p for this reaction at 25°C?

Solution

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Use the equation $K_p = K_c(RT)^{\Delta n}$. In this reaction, $\Delta n = 3 - 2 = 1$, so

$$K_p = K_c(RT)^{\Delta n} = (7.3\times10^{34})(0.0821\times298)^1 = 1.8\times10^{36}$$

2. 0.96 g of HI was heated to attain equilibrium $2HI \rightleftharpoons H_2 + I_2$. The reaction mixture on titration requires 15.7 mL of *N*/10 hyposolution. Calculate the degree of dissociation of HI.

Solution

Consider the reaction

	$2HI$	$\rightleftharpoons$	H_2	+	I_2
Initial moles	$0.96/128 = 7.5\times10^{-3}$		0		0
At equilibrium	$7.5\times10^{-3} - x$		$x/2$		$x/2$

Molarity of I_2 formed at equilibrium = molarity of hyposolution used

$$\frac{W}{M}\times1000 = 15.7\times\frac{1}{10}$$

Number of moles of I_2 formed at equilibrium = $\frac{1.57\times10^{-3}}{2} = 0.785\times10^{-3}$

$$\frac{x}{2} = 0.785\times10^{-3} \Rightarrow x = 1.57\times10^{-3}$$

Degree of dissociation = $\frac{x}{7.5}\times10^{-3} = \frac{1.57\times10^{-3}}{7.5\times10^{-3}} = 0.209 = 20.9\%$

3. N_2O_4 dissociates as $N_2O_4 \rightleftharpoons 2NO_2$. At 55°C and 1 atm, the percentage dissociation is 50.3%. At what pressure and same temperature, the equilibrium mixture will have the ratio of N_2O_4: NO_2 as 1:8?

Solution

Consider the reaction

	N_2O_4	$\rightleftharpoons$	$2NO_2$
Initial	1		0
At equilibrium	$1-\alpha$		2α

$$K_p = \frac{(2\alpha)^2}{1-\alpha}\cdot\left(\frac{p}{1+\alpha}\right)^{\Delta n}$$

$$K_p = \frac{4\alpha^2 p}{1-\alpha^2} = 4\times(0.503)^2\times1/1-(0.503)^2 = 1.355$$

	N_2O_4	$\rightleftharpoons$	$2NO_2$
Initial moles	1		0
At equilibrium	$1-x$		$2x$

Given that $N_2O_4/NO_2 = \frac{1-x}{2x} = \frac{1}{8} \Rightarrow x = 0.8$

$$K_p = \frac{4x^2\times p}{1-1.355} = \frac{4\times(0.8)^2\times p}{1-(0.8)^2} = 0.19 \text{ atm}$$

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4. One mole of each $Cl_2(g)$ and $Br_2(g)$ were taken in a 2 L flask, sealed and heated to some temperature, where the following equilibrium was established, and at equilibrium, the mol% of ClBr(g) was 40%.

$$Cl_2(g) + Br_2(g) \rightleftharpoons 2BrCl(g) \qquad (1)$$

At equilibrium, 2 mol of NO gas was added where the following additional equilibrium was established:

$$2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) \qquad (2)$$

At new equilibrium, mixture was found to contain 0.9 mol of BrCl(g). Determine K_C for the two equilibria in Eqs. (1) and (2).

Solution

For the reaction

	$Cl_2(g)$	$+ Br_2(g)$	$\rightleftharpoons 2BrCl(g)$
Initial	1 mol	1mol	0 mol
At equil	$1-x$	$1-x$	$2x$

Total number of moles at equilibrium = $1-x+1-x+2x = 2$

% mol of ClBr = $2x/2$= 40% =0.4, therefore, $x = 0.4$

The equilibrium constant can be represented as

$$K_c = \frac{[2x]^2}{(1-x)(1-x)} = \frac{4(0.4)^2}{0.6\times 0.6} = 1.78$$

5. At a certain temperature, $K_C = 4.50$ for the reaction

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

If 0.300 mol of N_2O_4 is placed into a 2.00 L container at that temperature, what will be the equilibrium concentrations of both gases?

Solution

The initial concentration of N_2O_4 is 0.300 mol/2.00 L = 0.150 M. As no NO_2 was placed in the reaction vessel, its initial concentration is 0.000 M. Let us take x to be the number of moles per liter of N_2O_4 that reacts, so the change in the N_2O_4 concentration is $-x$. Because of the stoichiometry of the reaction, the NO_2 concentration must increase by $2x$, so its change in concentration is $+2x$. Hence, the concentration table can be represented as

	$N_2O_4(g)$	$\rightleftharpoons 2NO_2(g)$
Initial concentration (mol L^{-1})	0.150	0.000
Changes in concentrations	$-x$	$+2x$
Concentration at equilibrium	$0.150-x$	$2x$

Now we substitute the equilibrium quantities into the expression for equilibrium constant.

$$\frac{[NO_2]^2}{[N_2O_4]} = 4.50 \Rightarrow \frac{(2x)^2}{(0.150-x)} = 4.50 \Rightarrow \frac{4x^2}{(0.150-x)} = 4.50$$

Note that where the changes in concentrations are unknown, it is convenient to let the coefficients of x be the same as the coefficients in the balanced equation. This ensures

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they are in the correct ratio. Because the equation involves terms in x^2, x, and a constant, we can use the quadratic formula to obtain the value of x.

$$4x^2 = 4.50(0.150 - x) = 0.675 - 4.50x$$

Arranging terms in the standard order gives

$$4x^2 + 4.50x - 0.675 = 0$$

Therefore, the quantities we will substitute into the quadratic formula are as follows: a = 4, b = 4.50, and c = – 0.675. Making these substitutions gives

$$x = \frac{-4.50 \pm \sqrt{(4.50)^2 - 4(4)(-0.675)}}{2(4)} = \frac{-4.50 \pm \sqrt{31.05}}{8} = \frac{-4.50 \pm 5.57}{8} \text{ M}$$

Because of the ± term, there are two values of x that satisfy the equation, x = 0.134 and x = –1.26. However, only the first value, x = 0.134, makes any sense chemically, as negative concentrations are not possible. Using this value, the equilibrium concentrations are

$$[N_2O_4] = 0.150 - 0.134 = 0.016 \text{ M}$$

$$[NO_2] = 2(0.134) = 0.268 \text{ M}$$

In general, whenever you use the quadratic equation in a chemical calculation, one root will be satisfactory and the other will lead to answers that do not make sense.

6. The standard Gibbs energy change for a reaction is ΔG^o= –41.8 kJ mol^{-1} at 700 K and 1 atm. Calculate the equilibrium constant of the reaction at 700 K.

Solution

The standard Gibbs energy and equilibrium constant are related by the equation:

$$\Delta G^o = -RT \ln K = -2.303\, RT \log K$$

Given that ΔG^o = –41.8 kJ = –41800 J, T = 700 K, and R = 8.314 J K^{-1} mol^{-1}. Substituting these values in the equation, we get

$$-41800 = -2.303 \times 8.314 \times 700 \log K$$

$$-41800 = -13403 \log K$$

$$\log K = \frac{41800}{13403} = 3.1187 \Rightarrow K = 1.312 \times 10^4$$

7. The reaction $NO_2(g) + NO(g) \rightleftharpoons N_2O(g) + O_2(g)$ has $\Delta G^o_{1273} = -9.67$ kJ. A 1.00 L reaction vessel at 1000°C contains 0.0200 mol NO_2, 0.040 mol NO, 0.015 mol N_2O, and 0.0350 mol O_2. Is the reaction at equilibrium? If not, in which direction will the reaction proceed to reach equilibrium?

Solution

$$Q_c = \frac{[N_2O][O_2]}{[NO_2][NO]} = \frac{(0.015/1)(0.0350/1)}{(0.200/1)(0.040/1)} = 0.0656$$

$$\Delta G^o = -2.303 RT \log K$$

$$-9670 = -2.303 \times 8.314 \times 1273 \log K$$

Hence, $\log K = 0.3967 \Rightarrow K = 2.49$

As $K > Q_C$, so the reaction will proceed in forward direction.

It is spontaneous as ΔG is negative.

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8. Four moles of hydrogen iodide are taken in a 10 L flask and kept at 800 K. It undergoes the following decomposition reaction until the equilibrium is attained. The equilibrium mixture is found to contain 0.42 mol of iodine. Calculate the equilibrium constant of the reaction.

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

Consider the reaction

	$2HI$	$\rightleftharpoons$	H_2	$+ I_2$
Initial no. of moles	4		0	0
At equilibrium	$4 - 2 \times 0.42$		0.42	0.42

$$K_C = \frac{[H_2][I_2]}{[HI]} = \frac{(0.42/10)(0}{(3.16/10)2} = 0.01767$$

9. Five grams of PCl_5 were completely vaporized at 250°C in a vessel of 1.9 L capacity. The mixture at equilibrium exerted a pressure of 1 atm. Calculate the degree of dissociation K_C and K_p for this reaction.

Solution

$$\text{Number of moles of } PCl_5 = \frac{\text{Weight}}{\text{Molecular weight}} = \frac{5}{208.5} = 0.024$$

The equilibrium for dissociation of $PC1_5$ may be represented as

	$PC1_5$	$\rightleftharpoons$	PCl_3	$+$	Cl_2
At $t = 0$	0.024		0		0
At equilibrium	$0.024 - \alpha$		α		α

Total moles of gas components = $0.024 - \alpha + 2\alpha = (0.024 + \alpha)$

We know that $pV = nRT$, so

$$1 \times 1.9 = (0.024 + \alpha)\ 0.0821 \times 523$$

$$(0.024 + \alpha) = \frac{1.9}{0.0821 \times 523} \Rightarrow \alpha = 0.0202$$

$$\text{Degree of dissociation, } \alpha = \frac{0.0202}{0.024} = 0.843$$

$$[PCl_5] = \frac{0.024 - \alpha}{1.9} = \frac{0.024 - 0.0202}{1.9} = \frac{0.0038}{1.9}$$

$$[PCl_3] = [Cl_2] = \frac{\alpha}{V} = \frac{0.0202}{1.9}$$

$$K_C = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{[0.0202/1.9]^2}{0.0038/1.9} = \frac{4.08 \times 10^{-4}}{1.9 \times 0.0038} = 0.0565 \text{ mol L}^{-1}$$

Now, $K_p = K_C\,(RT)^{\Delta n}$, where $K_C = 0.0565$ and $\Delta n = 2 - 1 = 1$

Solving, we get $K_p = 0.0565 \times (0.0821 \times 523) = 2.42$ 6<

10. The equilibrium constant for the reaction, $I_2(s) + Br_2(g) \rightleftharpoons 2IBr(g)$ is 0.164 at 25°C.

(a) Bromine gas is introduced into a container with excess solid iodine. The pressure and temperature are 0.164 atm and 25°C, respectively. Find the partial pressure of IBr formed at equilibrium. Assume that the vapor pressure of $I_2(s)$ is negligible.

(b) Calculate $\Delta G°$ for the above reaction.

Solution

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(a) $K_p = \frac{p_{IBr}^2}{p_{Br_2}} = 0.164$

Total equilibrium pressure $= 0.164\,\text{atm} = p_{IBr} + p_{Br_2}$

$$p_{Br_2} = 0.164 - p_{IBr}$$

Substituting in the above equation:

$$0.164 = \frac{p_{IBr}^2}{(0.164 - p_{IBr})}$$

$$p_{IBr}^2 + 0.164\ p_{IBr} - 0.0269 = 0$$

Solving, p_{IBr} = 0.1014 atm.

(b) $\Delta G^o = -RT\ln K = -8.314 \times 298 \times 2.303 \log 0.164 = 4.48$ kJ

11. Equilibrium constant, K_C, for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500 K is 0.061. At a particular time, the analysis shows that the composition of the reaction mixture is 3.00 mol L^{-1} N_2, 2.00 mol L^{-1} H_2, and 0.500 mol L^{-1} NH_3. Is the reaction at equilibrium?
If not, in which direction does the reaction tend to proceed to reach equilibrium?

Solution

Given that $[N_2]$ = 3.0 mol L^{-1}, $[H_2]$ = 2.0 mol L^{-1}, $[NH_3]$ = 0.50 mol L^{-1}

The reaction quotient for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3$ is

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(0.50)^2}{3\times(2)^3} = 0.0104$$

We have K_C = 0.061. Thus, at this stage of the reaction, $Q_C < K_C$. Therefore, the reaction is not at equilibrium. The reaction is on the left of the equilibrium state. That is, the reaction tends to move toward products side to form more products.

12. In the reaction $CuSO_4.3H_2O(s) \rightleftharpoons CuSO_4.H_2O(s)$; the dissociation pressure is 7×10^{-3} atm at 25°C and ΔH^o = 2750 cal. What will be the dissociation pressure at 127°C?

Solution

For given reaction, $K_p = (p_{H_2O})^2$

So, $K_p(25°C) = (7 \times 10^{-3})^2$ atm^2 = 4.9×10^{-5} atm^2

As ΔH^o = 2750 cal, so using Vant Hoff equation

$$\log\frac{[K_p(127°C)]}{[K_p(25°C)]} = \frac{\Delta H^o}{2.303\times R}\left[\frac{400-298}{400\times 298}\right] \Rightarrow \log\frac{[K_p(127°C)]}{[4.9\times10^{-5}]} = \frac{2750}{2.303\times2}\left[\frac{102}{119200}\right]$$

$K_p(127°C) = 3.2426 \times 4.9 \times 10^{-5}$

so P_{H_2O} at $127°C = \sqrt{K_p(127°C)} = 1.26\times10^{-2}$ atm

13. The vapor density of a mixture of NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the number of moles of NO_2 in 100 g mixture at equilibrium.

Solution

$N_2O_4 \rightleftharpoons 2NO_2$

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1 0
$1-\alpha$ 2α
Here x is degree of dissociation of N_2O_4

Normal vapor density of $N_2O_4 = \frac{28+64}{2} = 46$

so $\alpha = \frac{46-38.3}{(2-1)38.3} = 0.2$

Now, moles of N_2O_4 = (l – α) = 0.8 mol
Moles of NO_2 = 2x = 0.4 mol
Weight of N_2O_4 = 92 × 0.8 = 73.60 g
Weight of NO_2 = 0.4 × 46 = 18.4 g
So, weight of mixture = 92 g
In 92 g mixture weight of NO_2 = 18.4 g

In 100 g mixture weight of $NO_2 = \frac{18.4}{92} \times 100 = 20\,\text{g}$

So, moles of NO_2 in 100 g mixture $= \frac{20}{46} = 0.434$ mol.

14. At 500°C, the decomposition of water into hydrogen and oxygen,

$$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$$

has $K_c = 6.0 \times 10^{-28}$. How many moles of H_2 and O_2 are present at equilibrium in a 5.00 L reaction vessel at that temperature if the container originally held 0.015 mol H_2O?

Solution

Given that $K_C = \frac{[H_2]^2[O_2]}{[H_2O]^2} = 6.0 \times 10^{-28}$

	[H_2O]	**[H_2]**	**[O_2]**
I	0.015 mol/5.00L	–	–
C	$-2x$	$+2x$	$+x$
E	$3.0 \times 10^{-3} - 2x$	$+2x$	$+x$

$$K_C = \frac{[2x]^2[x]}{[3.0 \times 10^{-3} - 2x]^2} = 6.0 \times 10^{-28}$$

Assume $2x \ll 3.0 \times 10^{-3}$.

$$\frac{4x^3}{(3.0 \times 10^{-3})^2} = 6.0 \times 10^{-28} \Rightarrow x = 1.1 \times 10^{-11}$$

Hence, $[H_2] = 2.2 \times 10^{-11}$ M, $[O_2] = 1.1 \times 10^{-11}$ M .

15. At 460°C, the reaction

$$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$$

has K_C = 85.0. Suppose 0.100 mol of SO_2, 0.0600 mol of NO_2, 0.0800 mol of NO, and 0.120 mol of SO_3 are placed in a 10.0 L container at that temperature. What will be the concentrations of all the gases when the system reaches equilibrium?

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Solution

All concentrations are a factor of 10 less than the stated value due to the container size. As $Q < K_c$ for this reaction, the reaction proceeds as written, that is, from left to right.

	$[SO_2]$	$[NO_2]$	$[NO]$	$[SO_3]$
I	0.0100	0.00600	0.00800	0.0120
C	$-x$	$-x$	$+x$	$+x$
E	$0.0100 - x$	$0.00600 - x$	$0.00800 + x$	$0.0120 + x$

Substituting the above values for equilibrium concentrations into the mass action expression gives:

$$K_c = \frac{[NO][SO_3]}{[SO_2][NO_2]} = \frac{(0.00800+x)(0.0120+x)}{(0.0100-x)(0.00600-x)} = 85.0$$

$0.00510 - 1.36(x) + 85.0(x)^2 = 9.60 \times 10^{-5} + 0.0200(x) + (x)^2$

$84.0(x)^2 - 1.38(x) + 0.00500 = 0$

The quadratic equation is used with the following values: $a = 84.0$; $b = -1.38$; $c = 0.00500$. Upon solving for x, we find that only the negative root is sensible:

$x = 0.00542$ M.

$[SO_2] = 0.0100 - 0.00542 = 0.0046$ M

$[NO_2] = 0.00600 - 0.00542 = 0.00058$ M

$[NO] = 0.00800 + 0.00542 = 0.0134$ M

$[SO_3] = 0.0120 + 0.00542 = 0.0174$ M

16. One mole of nitrogen is mixed with 3 mol of hydrogen in a 4 L container. If 0.25% of nitrogen is converted to ammonia by the following reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Calculate the equilibrium constant (K_c) in concentration units. What will be the value of K_c for the following equilibrium?

$$\tfrac{1}{2}N_2(g) + \tfrac{3}{2}H_2(g) \rightleftharpoons NH_3(g)$$

Solution

Consider the reaction

	N_2	$+ 3H_2$	$\rightleftharpoons$	$2NH_3$
Initial no. of moles	1	3		0
At equilibrium	$1 - x$	$3 - 3x$		$2x$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(2x/V)^2}{(1-x/V)(3-3x/V)^3} = \frac{4\times 2V^2}{(1-x)(3-3x)^3}$$

We know that $x = 0.25/100$ and $V = 4$ L. Substituting, we get $K_C = 1.49 \times 10^{-5}$.

On multiplying the reaction by (1/2), we get

$$K'_C = (K_C)^{1/2} = (1.49 \times 10^{-5})^{1/2} = 3.86 \times 10^{-3}$$

17. Air containing 79% of nitrogen and 21% of oxygen by volume is heated at 2200 K and 1 atm until equilibrium is established according to the reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$. If the K_p of the reaction is 1.1×10^{-3}, calculate the amount of nitric oxide produced in terms of volume percent.

Solution

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$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

At equilibrium, we have $[N_2] = 0.79\ (1 - \alpha)$;

$[O_2] = 0.21(1 - \alpha)$, $[NO] = \alpha$

Total number of moles = $0.79\ (1 - \alpha) + 0.21\ (1 - \alpha) + 2\alpha = 1 + \alpha$

$$p_{N_2} = \frac{0.79(1-\alpha)}{1+\alpha} \times 1;\ \ p_{O_2} = \frac{0.21(1-\alpha)}{1+\alpha} \times 1;\ \ p_{NO} = \frac{2\alpha}{1+\alpha}.1$$

$$K_p = \frac{p_{NO}^2}{p_{N_2} \cdot p_{O_2}} = \frac{4\alpha^2}{0.79 \times 0.21(1-\alpha)^2} = 1.1 \times 10^{-3}$$

where $\alpha = 0.0067$ and vol. % of NO = 1.33.

18. N_2O_4 is 25% dissociated at 37°C and 1 atm pressure. Calculate:

(a) K_p for $N_2O_4 \rightleftharpoons 2NO_2$

(b) percent dissociation at 37°C and 0.1 atm.

Solution

(a) The reaction can be expressed as

$$N_2O_4 \rightleftharpoons 2NO_2$$

$$1 \qquad 0$$

$$1 - \alpha \qquad 2\alpha$$

The total number of moles = $1 - \alpha + 2\alpha = 1 + \alpha$.

Therefore, $p_{(N_2O_4)} = \left(\frac{1-\alpha}{1+\alpha}\right) \times p$ and $p_{(NO_2)} = \left(\frac{2\alpha}{1+\alpha}\right) \times p$

Given that $\alpha = 0.25$ and $p = 1$ atm. So,

$$p_{(N_2O_4)} = \left(\frac{1-0.25}{1+0.25}\right) = \frac{0.75}{1.25} = 0.6 \text{ atm and } p_{(NO_2)} = 0.4 \text{ atm}$$

$$K_p = \frac{(p_{(NO_2)})^2}{p_{(N_2O_4)}} = \frac{(0.4)^2}{(0.6)} = \frac{0.16}{0.6} = 0.267 \text{ atm}$$

(b) Suppose the degree of dissociation of N_2O_4 at 0.1 atm is x.

$$p_{(N_2O_4)} = \left(\frac{1-x}{1+x}\right) \times 0.1 \text{ and } p_{(NO_2)} = \left(\frac{2x}{1+x}\right) \times p$$

Hence, $$K_p = \frac{(p_{(NO_2)})^2}{p_{(N_2O_4)}} = \frac{\left(\frac{2x}{1+x}\right)^2 \times 0.1}{\left(\frac{1-x}{1+x}\right) \times 0.1} \Rightarrow 0.267 = \frac{4x^2}{1-x^2}$$

Solving, we get $x = 0.632 \Rightarrow 63.2\%$

19. A mixture of N_2 and H_2 in the ratio 1:3 at 50 atm and 650°C is allowed to react till equilibrium is reached. Ammonia present at equilibrium was at 25 atm pressure. Calculate the equilibrium constant for the reaction.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_2(g)$$

Solution

Equilibrium of formation of NH_3 may be given as,

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	N_2	+ $3H_2$	$\rightleftharpoons$ $2NH_3$
At $t = 0$	1	3	2
At equilibrium	$(1-x)$	$3(1-x)$	$2x$

Total moles at equilibrium = $1 - x + 3(1 - x) + 2x = 4 - 2x$

$$\text{Mole fraction of } NH_3 = \frac{2x}{(4-2x)}$$

$$\text{Partial pressure } NH_3 = \frac{2x}{(4-2x)} \times p = \frac{2x}{(4-2x)} \times 50$$

$$25 = \frac{2x \times 50}{(4-2x)} = 0.666$$

$$K_p = \frac{16x^2(2-x)^2}{27p^2(1-x)^4} = \frac{16 \times (0.666)^2 (2-0.666)^2}{27 \times (50)^2 \times (0.334)^2} = 1.677 \times 10^{-3}$$

20. At 473 K, equilibrium constant K_c for decomposition of phosphorus pentachloride PCl_5 is 8.3×10^{-3}. If decomposition is depicted as

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g); \quad \Delta_r H^o = 124.0 \text{ kJ mol}^{-1}$$

(a) What is the value of K_c for the reverse reaction at the same temperature?
(b) What would be the effect on K_c if (i) more PCl_5 is added, (ii) pressure is increased, or (iii) temperature is increased?

Solution

(a) Expression of K_c for the reaction is

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

Value of K_c for the reverse reaction at the same temperature is

$$K_c' = \frac{1}{K_c} = \frac{1}{8.3 \times 10^{-3}} = 1.2048 \times 10^2 = 120.48$$

(b) (i) K_c would remain the same because in this case, the temperature remains the same.
(ii) K_c is constant at constant temperature. Thus, in this case, K_c would not change.
(iii) In an endothermic reaction, the value of K_c increases with an increase in temperature. As the given reaction is an endothermic reaction, the value of K_c will increase if the temperature is increased.

21. At 540 K, 0.10 mol of PCl_5 is heated in an 8 L flask. The pressure of the equilibrium mixture is found to be 1.0 atm. Calculate K_p and K_c for reaction.

[IIT-JEE 1998]

Solution

Consider the reaction

	PCl_5	$\rightleftharpoons$ PCl_3	+ Cl_2
Initial moles	0.1	0	0
At equilibrium	$0.1 - x$	x	x

Total number of moles = $0.1 - x + x + x = 0.1 + x$

As $pV = nRT$, we have

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$$n = \frac{pV}{RT} = \frac{1 \times 8}{0.0821 \times 540} = 0.18$$

So, $0.1 + x = 0.18 \Rightarrow x = 0.08$

$$K_C = \frac{(0.08/8)(0.08/8)}{(0.02/8)} = 0.04$$

$$K_p = (K_c RT)^{\Delta n} = 0.04 \times (0.0821 \times 540)^1 = 1.77$$

22. The K value for the reaction: $H_2 + I_2 \rightleftharpoons 2HI$, at 460°C is 49. If the initial pressure of H_2 and I_2 is 0.5 atm, respectively, determine the partial pressure of each gas at equilibrium. **[IIT-JEE 1999]**

Solution

Consider the reaction

	H_2	+ I_2	$\rightleftharpoons$	2HI
Initial number of moles	0.5	0.5		0
At equilibrium	$0.5 - x$	$0.5 - x$		$2x$

$$K = \frac{(2x)^2}{(0.5 - x)^2} = 49 \Rightarrow 7 = \frac{2x}{0.5 - x} \Rightarrow x = 0.39$$

Pressure of H_2 and I_2 at equilibrium = 0.5 – 0.39 = 0.11 atm

Pressure of HI at equilibrium = 2 × 0.39 = 0.78 atm

23. Calculate the value of log K for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 25°C. The standard enthalpy of formation of $NH_3(g)$ is –46 kJ and standard entropies of $N_2(g)$, $H_2(g)$, and $NH_3(g)$ are 191, 130, 192 J $K^{-1}mol^{-1}$, respectively. (R = 8.314 J mol^{-1} K^{-1}) [IIT-JEE 1999]

Solution

Consider the reaction

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

From the given values, we have

$\Delta S^o = \Delta S^o(\text{products}) - \Delta S^o(\text{reactants})$

$= 2\Delta S^o(NH_3) - [\Delta S^o(N_2) + 3\Delta S^o(H_2)]$

$= 2 \times 192 - [191 + 3 \times 130] = -197$ J K^{-1} mol^{-1} = –0.197 kJ K^{-1} mol^{-1}

$\Delta G^o = \Delta H^o - T\Delta S^o = 2x(-46) - (298 \times -0.197) = -92 + 58.706 = -33294$ J

Now, $\Delta G^o = -2.303\, RT \log K$. So,

$$\log K = \frac{\Delta G^o}{2.303RT} = \frac{33294}{2.303 \times 8.314 \times 298} = 5.845$$

$\log K = \Delta G^o/2.303RT = 33294/2.303 \times 8.314 \times 298 = 5.845$

24. The value of K_p is 1×10^{-3} atm^{-1} at 25°C for the reaction: $2NO + Cl_2 \rightleftharpoons 2NOCl$. A flask contains NO at 0.02 atm and at 25°C. Calculate the moles of Cl_2 that must be added if 1% of the NO is to be converted to NOC1 at equilibrium. The volume of the

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flask is such that 0.2 mol of gas produces 1 atm pressure at 25°C. (Ignore probable association of NO toN_2O_2.) [IIT-JEE 2001]

Solution

Consider the reaction:

	2NO	+	Cl_2	$\rightleftharpoons$	2NOCl
Initial pressure	0.02		p		0
At equilibrium	(0.02 – 0.02/100)		(p – 0.01/100)		0.02/100
	198×10^{-4}		$p - 10^{-4}$		2×10^{-4}

$$K_p = \frac{p_{\text{NOCl}}^2}{p_{\text{NO}}^2 \times p_{\text{Cl}_2}} \Rightarrow 10^{-3} = \frac{(2\times10^{-4})^2}{(198\times10^{-4})^2 \times (p-10^{-4})^2} \Rightarrow p = 0.1021 \text{ atm}$$

We know that $pV = nRT$, so

$$V = \frac{nRT}{p} = \frac{0.2\times0.0821\times273}{1} = 4.88\,\text{L}$$

$$n = \frac{pV}{RT} = \frac{0.1021\times4.887}{0.0821\times298} = 0.0204 \text{ mol}$$

25. The equilibrium constant K_c for the reaction $N_2O_4 \rightleftharpoons 2NO_2$ in chloroform at 291 K is 1.14. Calculate the free energy change of the reaction when the concentration of the two gases is 0.5 mol dm^{-3} each at the same temperature. (R = 0.082 L atm K^{-1}mol^{-1})

Solution

From the given data, we have T = 291 K; R = 0.082 L atm K^{-1}mol^{-1}, K_c = 1.14; C = 0.5 mol dm^{-3}.

As $Q_p = Q_c\,(RT)^{\Delta n}$ and Δn = 2 – 1 = 1 in this case, the reaction quotient Q_c for the reaction is

$$Q_c = \frac{0.5}{(0.082\times291)} = 11.93$$

The equilibrium constant is

$$K_p = K_c(RT)^{\Delta n} = 1.14\times(0.082\times291) = 27.1$$

Substituting these values in the equation

$$\Delta G = \Delta G^o + RT\ln Q_p = -RT\ln K_p + RT\ln Q_p = -2.303RT(\log K_p - \log Q_p)$$

we get

$$\Delta G = -(0.082\times291\times2.303)[\log 27.2 - \log 11.93]$$
$$= -54.95(1.4346 - 1.0766) = -19.67\,\text{L atm}$$

Additional Objective Questions

Single Correct Choice Type

1. Equilibrium constants are given (in atm) for the following reactions at 0°C

$$SrCl_2\cdot6H_2O(s) \rightleftharpoons SrCl_2\cdot2H_2O(s) + 4H_2O(g); \quad K_p = 6.89\times10^{-12}$$

The vapor pressure of water at 0°C is 4.58 torr. Calculate the pressure of water vapor in equilibrium at 0°C

(A) 2.6×10^{-3} atm (B) 1.62×10^{-3} atm

(C) 1.20×10^{-4} atm (D) 4.28×10^{-3} atm

Solution

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(B) $K_p = p_{H_2O}^4 = 6.89 \times 10^{-12}$ therefore $p_{H_2O} = 1.62 \times 10^{-3}$ atm

2. Predict which of the following reaction will have appreciable concentration of reactants and products:

(A) $Cl_2(g) \rightleftharpoons 2Cl(g); \quad K_c = 5 \times 10^{-39}$

(B) $Cl_2(g) + 2NO(g) \rightleftharpoons 2NOCl(g); \quad K_c = 3.7 \times 10^8$

(C) $Cl_2(g) + 2NO_2(g) \rightleftharpoons 2NO_2Cl(g); \quad K_c = 1.8$

(D) All of these

Solution

(C) As K_c is neither too high nor very low, reactants and products will be present in comparable amounts.

3. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ When 46 g I_2 and 1 g of H_2 are heated, an equilibrium is reached at 450°C. The equilibrium mixture contained 1.9 g of I_2. How many moles of I_2 and HI are present at equilibrium?

(A) 0.0075 and 0.147 mol (B) 0.0050 and 0.147 mol

(C) 0.0075 and 0.347 mol (D) 0.0052 and 0.347 mol

Solution

(C) Moles of I_2 taken $= \frac{46}{254} = 0.181$

Moles of H_2 taken $= \frac{1}{2} = 0.5$

Moles of I_2 remaining $= \frac{1.9}{254} = 0.0075$

Moles of I_2 used = 0.181 – 0.0075 = 0.1735

Moles of H_2 used = 0.1735

Moles of H_2 remaining = 0.5 – 0.1735 = 0.3265

Moles of HI formed = 0.1735 × 2 = 0.347

4. The partial pressures of CH_3OH, CO, and H_2 in the equilibrium mixture for the reaction $CO + 2H_2 \rightleftharpoons CH_3OH$ at 427°C are 2.0, 1.0, and 0.1 atm, respectively. The value of K_p for the decomposition of CH_3OH into CO and H_2 is

(A) 1×10^2 atm (B) 2×10^2 atm^{-1}

(C) 50 atm^2 (D) 5×10^{-3} atm^2

Solution

(B) $K_p = \frac{p_{CH_3OH}}{p_{CO} \times p_{H_2}^2} = \frac{2.0}{1.0 \times (0.1)^2} = 200$ atm^{-1}

5. Consider the following equilibrium in a closed container $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. At a fixed temperature, the volume of the reaction container is halved. For this change, which of the following statements holds true regarding the equilibrium constant (K) and degree of dissociation (α): (IIT-JEE 2002)

(A) Neither K_p nor α changes (B) Both K and α change

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(C) K_p changes but α does not (D) K does not change but α changes

Solution

(D)

6. For the reaction $2NO_2(g) + O_2(g) \rightleftharpoons N_2O_5(g$

If the equilibrium constant is K_p, then the equilibrium constant for the reaction $2N_2O_5(g) \rightleftharpoons 4NO_2 + O_2(g)$ would be

(A) K_p^2 (B) $\frac{2}{K_p}$ (C) $\frac{1}{K_p^2}$ (D) $\frac{1}{\sqrt{K_p}}$

Solution

(C) By writing the equation for the reverse reaction

$$N_2O_5(g) \rightarrow 2NO_2(g) + \tfrac{1}{2}O_2(g)$$

as $\frac{1}{K_p}$ and relating it to the desired reaction, the result can be obtained.

7. A container whose volume is V contains an equilibrium mixture that consists of 2 mol each of PCl_5, PCl_3, and Cl_2 (all gases). The pressure is 30.3975 kPa and temperature is T. A certain amount of $Cl_2(g)$ is now introduced keeping the pressure and temperature constant until the equilibrium volume changes to $2V$. Calculate the amount of Cl_2 that was added

(A) 10/3 mol (B) 20/3 mol (C) 1/3 mol (D)1/4 mol

Solution

(B) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

At equilibrium, 2 mol 2 mol 2 mol

Total pressure = 30.3975 kPa = 3 atm = p(say)

$$K_p = \frac{p_{PCl_3} \times p_{Cl_2}}{p_{PCl_5}} = \frac{p/3 \times p/3}{p/3} = p/3 \qquad (1)$$

When chlorine is added to the system, the system will try to nullify the effect, and hence, formation of PCl_5 will be preferred.

As p and T are constant, and given that $V_2 = 2V_1$; $n_1 = 6$ mol, so

$$\frac{V_1}{V_2} = \frac{n_1}{n_2} \Rightarrow n_2 = 12 \text{ mol}$$

Say, "a" moles of Cl_2 were added

	PCl_5 $\rightleftharpoons$	PCl_3 +	Cl_2
Initial moles	2	2	2
Final moles	$2 + x$	$2 - x$	$2 + a - x$

Total number of moles = 12 = $6 + a - x \Rightarrow a - x = 6$

$$K_p = \frac{\frac{2+a-x}{12} \times \frac{2-x}{12}}{\frac{2+x}{12}} \times p = \frac{(2+a-x)(2-x)}{12 \times (2+x)} p = \frac{(2+6)(2-x)}{12 \times (2+x)} p = \frac{2}{3} \times \frac{(2-x)}{(2+x)} p \qquad (2)$$

Equating Eqs. (1) and (2), we get x =2/3 and hence a = 20/3 mol.

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Hence,20/3 mol of Cl_2 was added.

8. For a certain equilibrium, over the temperature range 500–700 K, the equilibrium constant, K_p confirms to the equation $\log K_p = 30.1 - \frac{6.36\times10^3}{T} - 5.7 \log T$

Calculate ΔG^o at 600 K.

(A) –16.04 kcal mol^{-1} (B) –14.60 kcal mol^{-1}

(C) –10.064 kcal mol^{-1} (D) –14.00 kcal mol^{-1}

Solution

(C) $\log K_p = 30.0 - \frac{6.36\times10^3}{T} - 5.7\log T$

$$\Delta G^o = -RT \ln K_p = -2.303RT \log K_p$$

$$= -2.303\times1.987\times600\left[30.1 - \frac{6.36\times10^3}{600} - 5.7\log 600\right]$$

$$= -2.303 \times 1.987 \times 600[30.1 - 10.6 - 5.7 \times 2.778]$$

$$= -[2.303 \times 1.987 \times 600 \times 3.6654] \text{ cal mol}^{-1}$$

$$= -10063.9 \text{ cal mol}^{-1} = -10.064 \text{ kcal mol}^{-1}$$

9. PCl_5 is 80% dissociated at 523 K. Calculate the vapor density of the equilibrium mixture at 523 K.

(A) 75.9 (B) 57.9 (C) 97.5 (D) 95.7

Solution

(B)Vapor density $= \frac{\text{Molecular weight of } PCl_5}{2} = \frac{208.5}{2} = 104.25$

Applying the formula,

$$0.8 = \frac{104.25 - d}{(2-1)d} \Rightarrow 0.8d = 104.25 - d \Rightarrow 1.8d = 104.25 \Rightarrow d = \frac{104.25}{1.8} = 57.9$$

10. In the melting of ice, which one of the conditions will be more favorable?

(A) High temperature and high pressure.

(B) Low temperature and low pressure.

(C) Low temperature and high pressure.

(D) High temperature and low pressure.

Solution

(A) Melting of ice is the endothermic process, and on melting, volume decreases. Hence, high temperature and high pressure is favorable.

11. Example of physical equilibrium is

(A) $H_2 + I_2 \rightleftharpoons 2HI$ (B) $CaCO_3 \rightleftharpoons CaO + CO_2$

(C) $H_2O(s) \rightleftharpoons H_2O(l)$ (D) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$

Solution

(C) Physical equilibrium, do not include any chemical change.

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12. For the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3$

If the partial pressure of SO_2, O_2, and SO_3 are 0.559, 0.101, and 0.331 atm, respectively. What would be the partial pressure of O_2 gas to get equal moles of SO_2 and SO_3.

(A) 0.188 atm (B) 0.288 atm (C) 0.388 atm (D) 0.488 atm

Solution

(B) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

$$K_p = \frac{[p_{SO_3}]^2}{[p_{SO_2}]^2[p_{O_2}]} = \frac{(0.331)^2}{(0.559)^2(0.101)} = 3.47$$

SO_2 and SO_3 have same number of moles; their partial pressure will be equal $p_{SO_3} = p_{SO_2}$, therefore $p_{SO_2}\frac{1}{3.47} = 0.288\,\text{atm}$

13. When ethanol and acetic acid were mixed together in equimolecular proportion, 66.6% is converted into ethyl acetate. Calculate K_c. Also calculate quantity of ester produced if 1 mol of acetic acid is treated with 0.5 and 4 mol of alcohol, respectively.

(A) 4, 0.93, 0.43 (B) 0.93, 4, 0.43

(C) 0.43, 0.93, 4 (D) 4, 0.43, 0.93

Solution

(D) $CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$

1	1	0	0
1 – 0.666	1 – 0.666	0.666	0.666

$$K_C = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]} = \frac{[0.666][0.666]}{[0.333][0.333]}$$

(A) Let x mol of ester is formed from 1 mol of acid and 0.5 mol of alcohol, then

$$K_C = \frac{x^2}{(1-x)(0.5-x)} = 4 \Rightarrow x = 0.43$$

(B) $K_C = \frac{x^2}{(1-x)(4-x)}$ or $4 = \frac{x^2}{(1-x)(4-x)} \Rightarrow x = 0.93$

14. Calculate the temperature at which the product gases at equilibrium at 1 atm will contain 400 ppm (parts per million) of carbon monoxide.

(A) 399 K (B) 400 K (C) 275 K (D) 200 K

Solution

(A) For the given reaction $p_{CO} \gg p_{CO_2}$, hence $K_p = \frac{1}{p_{CO}} = \frac{1}{400 \times 10^6} = 2500$

$$\Delta G^o = -RT \ln K_p;\ \ln K_p = \frac{\Delta G^o}{RT} = \frac{20700 + 11.97T}{RT}$$

The equation when solved for T using $R = 8.314\ K^{-1}\ mol^{-1}$ gives $T = 399$ K.

15. For the reversible reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500°C, the value of K_p is 1.44×10^{-5} when partial pressure is measured in atmospheres. The corresponding value of K_C, with concentration in mol L^{-1}, is (IIT-JEE 2000)

(A) $\frac{1.44 \times 10^{-5}}{(0.082 \times 500)^{-2}}$ (B) $\frac{1.44 \times 10^{-5}}{(8.314 \times 773)^{-2}}$

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(C) $\frac{1.44\times10^{-5}}{(0.082\times773)^2}$ (D) $\frac{1.44\times10^{-5}}{(0.082\times773)^{-2}}$

Solution

(B) $K_p = K_C\,(RT)^{\Delta n}$ where $K_p = 1.44 \times 10^{-5}$, $\Delta n = 2 - 4 = -2$. Hence,

$$K_C = \frac{1.44\times10^{-5}}{(8.314\times773)^{-2}}$$

16. Le Chatelier's principle is applicable to

(A) only homogeneous chemical reversible reaction.

(B) only heterogeneous chemical reversible reaction.

(C) only physical equilibria.

(D) all systems chemical or physical, in equilibrium.

Solution

(D)

17. At temperature T, a compound $AB_2(g)$ dissociates according to the reaction $2AB_2(g) \rightleftharpoons 2AB(g) + B_2(g)$ with degree of dissociation α, which is small compared with unity. The expression for K_p in terms of α and the total pressure p_T is

(A) $\frac{p_T\alpha^3}{2}$ (B) $\frac{p_T\alpha^2}{3}$ (C) $\frac{p_T\alpha^2}{2}$ (D) None of these

Solution

(A) For the given equilibrium, the equilibrium concentration is

$$2AB_2(g) \rightleftharpoons 2AB(g) + B_2(g)$$

Equilibrium concentration $c(1-\alpha)$, $c\alpha$, $\frac{c\alpha}{2}$

Therefore, $$K_p = \frac{(p_{B_2})(p_{AB})^2}{(p_{AB_2})^2} = \frac{(c\alpha/2)\times(c\alpha)^2}{[c(1-\alpha)]^2}\times\frac{p_T}{[c(1+\alpha/2)]} = \frac{\alpha^3\times p_T}{2(1-\alpha)^2\left(1+\frac{\alpha}{2}\right)}$$

As a is small compared with unity, so $1-\alpha \simeq 1$ and $1+\frac{\alpha}{2} \simeq 1$ therefore $K_p = \frac{\alpha^3\times p_T}{2}$

18. Starting with 3:1 mixture of H_2 and N_2 at 450°C, the equilibrium mixture is found to be 9.6% NH_3, 22.6% N_2, and 67.8% H_2 by volume. The total pressure is 50 atm. What will be the value of K_p. The reaction is $N_2 + 3H_2 \rightleftharpoons 2NH_3$

(A) $3.25 \times 10^{-5}\,atm^{-2}$ (B) $5.23 \times 10^{-5}\,atm^{-2}$

(C) $6.23 \times 10^{-2}\,atm^{-2}$ (D) $8 \times 10^{-5}\,atm^{-2}$

Solution

(B)The ratio of number of moles will be the same as the ratio of volume.

According to Dalton's law, the partial pressure of a gas in a mixture is given by the product of its volume fraction and the total pressure.

Therefore, the equilibrium pressure of each gas is,

$$p_{NH_3} = \frac{9.6}{100}\times50\,atm = 4.8\,atm$$

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$$p_{N_2} = \frac{22.6}{100} \times 50\,\text{atm} = 11.3\,\text{atm}$$

$$p_{H_2} = 67.6/100 \times 50\,\text{atm} = 33.9\,\text{atm}$$

$$\text{Total pressure} = 50\,\text{atm}$$

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

Substituting the values of partial pressures,

$$K_p = \frac{(4.80\,\text{atm})^2}{(11.3\,\text{atm})(33.9\,\text{atm})^3} = 5.23 \times 10^{-5}\ \text{atm}^{-2}$$

19. Find out the value of K_c for each of the following equilibrium from the value of K_p:

(i) $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g);\quad K_p = 1.8 \times 10^{-2}$ at 500 K

(ii) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g);\quad K_p = 167$ at 1073 K

the value of K_C are, respectively

(A) 4.4×10^{-4} and 1.90 (B) 8.8×10^{-4} and 3.8

(C) 4.4×10^{-4} and 1.90 (C) 8.8×10^{4} and 3.8

Solution

(A) (i) $K_p = K_C(RT)^{\Delta n}$

(ii) Therefore, $\Delta n = 1$ and

$$K_c = \frac{K_p}{RT} = \frac{167}{0.0821 \times 1073} = 1.90$$

Since $\Delta n = 1$, therefore $K_c = \dfrac{K_p}{RT} = \dfrac{1.8 \times 10^{-2}}{0.082 \times 500} = 4.4 \times 10^{-4}$

20. At constant temperature, the equilibrium constant (K) for the decomposition reaction $N_2O_4(\alpha/2) \rightleftharpoons 2NO_2$ is expressed by $K_p = \dfrac{4x^2 p}{(1-x^2)}$ where p is the pressure and x is the extent of decomposition. Which one of the following statements is true? (IIT-JEE 2000)

(A) K increases with increase of p

(B) K increases with increase of x

(C) K increases with decrease of x

(D) K remains constant with change in p and x

Solution

(D)

21. The vapor density of N_2O_4 at a certain temperature is 30. The percentage dissociation of N_2O_4 at this temperature is

(A) 55.5% (B) 60% (C) 70% (D) 53.3%

Solution

(D) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$, molecular mass of N_2O_4 = (28 + 64) = 92

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Vapor density, $D = \frac{92}{6} = 46$

Let the degree of dissociation be x. Applying the relationship,

$$x = \frac{D-d}{d} = \frac{(46-30)}{30} = \frac{16}{30} = 0.533$$

Degree of dissociation = 53.3%.

22. The equilibrium constant of the reaction $A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$ at 100°C is 50. If 1 L flask containing 1 mol of A_2 is connected to a 2 L flask containing 2 mol of B_2, the number of moles of AB formed at 373 K will be

(A) 1.886 (B) 2.317 (C) 0.943 (D) 18.86

Solution

(A) The equilibrium is represented as:

	$A_2(g)$ +	$B_2(g)$ $\rightleftharpoons$	$2AB(g)$
Initial concentration	1	2	0
Moles at equilibrium	$1 - x$	$2 - x$	$2x$

Total volume = 1 + 2 = 3 L

$$[A_2] = \frac{1-x}{3}, [B_2] = \frac{2-x}{3}, [AB] = \frac{2x}{3}$$

$$K = \frac{[AB]^2}{[A_2][B_2]} = \frac{2x/3}{(1-x)/3(2-x)/3} = 50$$

On solving, we get $23x^2 - 75x + 50 = 0 \Rightarrow x = 2.31$ or 0.943.

As x cannot be more than 1, so $x = 0.943$

Number of moles of AB formed = 2 × 0.943 = 1.886 mol.

23. NO and Br_2 at initial partial pressure of 98.4 and 41.3 torr, respectively, were allowed to react at 300 K. At equilibrium, the total pressure was 110.5 torr. Calculate the value of K_p for the following reaction at 300 K.

$2NO(g) + Br(g) \rightleftharpoons 2NOBr(g)$

(A) 124 (B) 134 (C) 154 (D) 174

Solution

(B) The given reaction is,

	$2NO(g)$ +	$Br_2(g)$ $\rightleftharpoons$	$2NHOBr(g)$
Initial	98.4	41.3	0
At equilibrium	$98.4 - x$	$41.3 - \frac{x}{2}$	x

Total pressure at equilibrium = 110.5 torr.

$$(98.4 - x) + \left(41.3 - \frac{x}{2}\right) + x = 110.5$$

Solving, we get $x = 58.4$ torr (760 torr = 1 atm)

$p_{NOBr} = 58.4$ torr $= 7.68 \times 10^{-2}$ atm.

$p_{NO} = 98.4 - 58.4 = 40$ torr $= 5.26 \times 10^{-2}$ atm.

$$K_p = \frac{p_{NOBr}^2}{p_{NO}^2 \times p_{Br_2}} = \frac{(7.68\times10^{-2})^2}{(5.26\times10^{-2})^2(1.59\times10^{-2})} = 134$$

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24. In the equilibrium $SO_2Cl_2 \rightleftharpoons SO_2 + Cl_2$ at 2000 K and 10 atm pressure, percent Cl_2 = percent SO_2 = 40 by volume. Then

(A) K_p= 2 atm (B) $\frac{n(SO_2Cl_2)}{n(SO_2)} = \frac{1}{4}$ at equilibrium

(C) K_p = 8 atm (D) $n(SOCl_2) = n(SO_2) = n(Cl_2)$

Solution

(C) At 40% SO_2 and 40% CO_2 by volume $p_{CO_2} = \left(\frac{40}{100}\right)p$ and $p_{SO_2} = \left(\frac{40}{100}\right)p$

So, $p_{SO_2Cl_2} = \left(\frac{20}{100}\right)p$ and given that p = 10 atm. Therefore,

Therefore, $K_p = \frac{(p_{CO_2})(p_{SO_2})}{(p_{SO_2Cl_2})} = \frac{(40/100)p \times (40/100)p}{(20/100)p} = \frac{16 \times 10}{2 \times 10} = 8$ atm

25. At a certain temperature, the equilibrium constant K_c is 16 for the reaction, $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$. If 1.0 mol each of all the four gases is taken in a 1 L container, the concentration of NO_2 at equilibrium would be

(A) 1.6 mol L^{-1} (B) 0.8 mol L^{-1} (C) 0.4 mol L^{-1} (D) 0.6 mol L^{-1}

Solution

(C) The reaction is

	$SO_2(g)$	$+ NO_2(g)$	$\rightleftharpoons SO_3(g)$	$+ NO(g)$
Initial concentration	1	1	1	1
At equilibrium	$1-x$	$1-x$	$1+x$	$1+x$

$$K = \frac{[SO_3][NO]}{[SO_2][NO_2]} = \frac{(1+x)(1+x)}{(1-x)(1-x)}$$

$$16 = \frac{(1+x)^2}{(1-x)} \Rightarrow \frac{(1+x)}{(1-x)} = 4 \quad \text{or} \quad x = 0.6$$

$[NO_2] = 1 - x = 1 - 0.6 = 0.4\,\text{mol}\,L^{-1}$

26. In a system $A(s) \rightleftharpoons 2B(g) + 3C(g)$. If the concentration of C at equilibrium is increased by a factor of 2, it will cause the equilibrium concentration of B to change to

(A) two times of its original value. (B) one half of its original value.

(C) $2\sqrt{2}$ times of its original value. (D) $1/2\sqrt{2}$ times of its original value.

Solution

(D) $\underset{1-x}{A(s)} \rightleftharpoons \underset{2x}{2B(g)} + \underset{3x}{3C(g)}$

$K_c = 4x^2 \times 27x^3 = 108x^5$

$A(s) \rightleftharpoons \underset{2x \times c}{2B(g)} + \underset{6x}{3C(g)}$

$108x^5 = 4x^2 \times c^2 \times 216x^3 \Rightarrow 108 = c^2 \times 216 \times 4$

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$$c^2 = \frac{108}{4 \times 216} \Rightarrow c = \frac{1}{2\sqrt{2}}$$

27. For the chemical reaction $3X(g) + Y(g) \rightleftharpoons X_3Y(g)$, the amount of X_3Y at equilibrium is affected by

(A) temperature and pressure. (B) temperature only.
(C) pressure only. (D) temperature, pressure, and catalyst.

[IIT-JEE 2009]

Solution

(A)

28. In reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, the concentration of H_2, I_2, and HI at equilibrium are 8.0, 3.0, and 28 mol L^{-1}, respectively. What will be the equilibrium constant?

(A) 30.61 (B) 32.66 (C) 29.40 (D) 20.90

Solution

(B) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

Applying law of mass action,

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

Given that $[H_2]$ = 8.0 mol L^{-1}, $[I_2]$ = 3.0 mol L^{-1} [HI]=28.0 mol L^{-1}

$$K_c = \frac{(28.0)^2}{(8.0) \times (3.0)} = 32.66$$

29. When two reactants A and B are mixed to give products C and D, the reaction quotient at the initial stages of the reaction (IIT-JEE 2000)

(A) is zero. (B) decreases with time.
(C) is independent of time. (D) increases with time.

Solution

(D) The equilibrium constant K for the reaction

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ is 1.64×10^{-4} atm at 400°C.

30. What will be the equilibrium constant at 500°C if heat of reaction in this temperature range is –105185.8 J.

(A) 0.144×10^{-4}atm (B) 0.144×10^{-3}atm
(C) 0.144×10^{-2}atm (D) 0.144×10^{-1}atm

Solution

(A) Given that $K_p = 1.64 \times 10^{-4}$atm, K_{p2} =?; T_1 = 400 + 273 = 673 K, T_2 = 500 + 273 = 773 K; ΔH = –105185.8 J, R = 8.314 J K^{-1} mol^{-1}

Applying the equation

$$\log K_{p_2} - \log K_{p_1} = \frac{\Delta H}{2.303R}\left(\frac{T_2 - T_1}{T_1 T_2}\right)$$

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$$\log K_{p_2} = \log 1.64 \times 10^{-4} - \frac{105185.8}{2.303 \times 8.314}\left(\frac{773-673}{773 \times 673}\right)$$

or $K_{p2} = 0.144 \times 10^{-4}$ atm.

31. The following reactions are known to occur in the body, $CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^{-1}$. If CO_2 escapes from the system, then

(A) pH will decrease.
(B) hydrogen ion concentration will diminish.
(C) H_2CO_3 concentration will remain unchanged.
(D) the forward reaction will be favored.

Solution

(B) This is according to Le Chatelier's principle.

32. For the reaction: $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$, the degree of dissociation (α) of HI(g) is related to equilibrium constant K_p by the expression

(A) $\frac{1+2\sqrt{K_p}}{2}$ (B) $\sqrt{\frac{1+2K_p}{2}}$ (C) $\sqrt{\frac{2K_p}{1+2K_p}}$ (D) $\frac{2\sqrt{2K_p}}{1+2\sqrt{K_p}}$

Solution

(D) $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$

$1-\alpha \qquad \frac{\alpha}{2} \qquad \frac{\alpha}{2}$

$$K_p = \frac{\left(\frac{\alpha}{2}p_T\right)^2}{(1-\alpha)^2 p_T^2} \text{ or } \frac{\alpha}{1-\alpha} = 2\sqrt{K_p} \text{ or } \alpha = \frac{2\sqrt{K_p}}{1+2\sqrt{K_p}}$$

33. For which of the following reactions the degree of dissociation cannot be calculated from the vapor density data

(I) $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$ (II) $2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$

(III) $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$ (IV) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

(A) I and III (B) II and IV (C) I and II (D) III and IV

Solution

(A) The degree of dissociation cannot be calculated from the vapor density data because here, the number of moles remains unchanged before and after reaching equilibrium.

34. An example of reversible reaction is

(A) $Pb(NO_3)_2 + 2NaI = PbI_2 + 2NaNO_3$ (B) $AgNO_3 + HCl = AgCl + HNO_3$
(C) $2Na + 2H_2O = 2NaOH + H_2$ (D) $KNO_3 + NaCl = KCl + NaNO_3$

Solution

(D)

35. For the reaction $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, at a given temperature, the equilibrium amount of CO_2 (g) can be increased by (IIT-JEE 1998)

(A) adding a suitable catalyst.

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(B) adding an inert gas.
(C) decreasing the volume of the container.
(D) increasing the amount of CO(g).

Solution

(D)

36. A reaction mixture containing H_2, N_2, and NH_3 has partial pressures 2, 1, and 3 atm, respectively, at 725 k. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 4.28×10^{-5} atm^{-2} at 725 K. In which direction the net reaction will go?

(A) Forward. (B) Backward.
(C) No net reaction. (D) Cannot be predicted.

Solution

(B) $Q_p = \dfrac{(p_{NH_3})^2}{p_{N_2} \times (p_{H_2})^3} = \dfrac{(3)^2}{(1)(2)^3} = \dfrac{9}{8}\ \text{atm}^{-1} = 1.125\ \text{atm}^{-2}$

37. Ammonium hydrogen sulphide dissociates according to the equation $NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$. If the observed pressure of the mixture is 1.12 atm at 106°C, what is the equilibrium constant K_p of the reaction?

(A) 0.3136 atm^2 (B) 2 atm^2
(C) 1 atm^2 (D) 0.4842 atm^2

Solution

(A) The reaction is

$$\begin{array}{lcc} & NH_4HS(s) \rightleftharpoons NH_3(g) & + H_2S(g) \\ t = 0 & 0 & 0 \\ t = t_{eq} & x\,\text{atm} & x\,\text{atm} \end{array}$$

Now, $p_T = 2x = 1.12 \Rightarrow x = 0.56;\ K_p = x^2 = 0.3136\,\text{atm}^2$

38. When sulphur in the form of S_8 is heated at 900 K, the initial pressure of 1 atm falls by 29% at equilibrium. This is because of conversion of some S_8 to S_2. Find the value of equilibrium constant for this reaction

(A) 1.16 atm^3 (B) 0.71 atm^3 (C) 2.55 atm^3 (D) 5.1 atm^3

Solution

(C) The reaction is

$$\begin{array}{lcc} & S_8(g) \rightleftharpoons & 4S_2(g) \\ \text{At } t = 0 & 1 & 0 \\ \text{At equilibrium} & 1 - 0.29 = 0.71\,\text{atm} & 4 \times 0.29 = 1.16\,\text{atm} \end{array}$$

$$K_p = \frac{[p_{S_2}]^4}{[p_{S_8}]} = \frac{(1.16)^4}{(0.71)} = 2.55\,\text{atm}^3$$

39. $BiCl_3(g) + H_2O(g) \rightleftharpoons BiOCl(s) + 2HCl(g)$ in the above system at equilibrium, the pressure of water vapor is doubled. The new equilibrium partial pressure of HCl increases

(A) 1.73 times (B) 1.414 times (C) 2.0 times (D) 1.0 times

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Solution

(B) When the denominator is increased by 2, $p_{HCl}^2 = 2$

Therefore $p_{HCl} = \sqrt{2} = 1.414$

40. K_1 and K_2 are equilibrium constant for reactions

(I) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ (II) $NO(g) \rightleftharpoons \frac{1}{2}N_2(g) + \frac{1}{2}O_2(g)$

Then, from (I) and (II)

(A) $K_I = \left(\frac{1}{K_{II}}\right)^2$ (B) $K_I = K_{II}^2$ (C) $K_I = \frac{1}{K_{II}}$ (D) $K_I = (K_{II})^0$

Solution

(A) $N_2 + O_2 \rightleftharpoons 2NO$ ...K_1

$2NO \rightleftharpoons N_2 + O_2$ ...$\frac{1}{K_1}$

$NO \rightleftharpoons \frac{1}{2}N_2 + O_2$ ...$K_2 = \left(\frac{1}{K_1}\right)^{1/2}$

$$K_2 = \left(\frac{1}{K_1}\right)^{1/2} \Rightarrow K_2^2 = \left(\frac{1}{K_1}\right) \Rightarrow K_1 = \left(\frac{1}{K_2^2}\right) \text{ or}$$

41. If a mixture of 0.482 mol N_2 and 0.933 mol of O_2 is placed in a 10 L reaction vessel and allowed to form N_2O at a temperature for which $K_c = 2.0 \times 10^{-37}$, then the composition of equilibrium mixture is

(A) N_2 = 0.482 mol L^{-1}, O_2 = 0.0933 mol L^{-1}, $N_2O = 6.6 \times 10^{-21}$ mol L^{-1}.

(B) N_2 = 0.482 mol L^{-1}, O_2 = 0.933 mol L^{-1}, $N_2O = 6.6 \times 10^{-21}$ mol L^{-1}.

(C) N_2 = 0.0482 mol L^{-1}, O_2 = 0.0933 mol L^{-1}, $N_2O = 6.6 \times 10^{-21}$ mol L^{-1}.

(D) N_2 = 0.0482 mol L^{-1}, O_2 = 0.0931 mol L^{-1}, $N_2O = 8.8 \times 10^{-21}$ mol L^{-1}.

Solution

(C) The reaction involved is

	$2N_2(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2N_2O(g)$
$t = 0$	0.482 mol		0.933 mol		0
Charge	$-2x$		$-x$		$+2x$
$t = t_{eq}$	$(0.0482 - 2x)$		$(0.0933 - x)$		$2x$

$$K_c = 2 \times 10^{-37} = \frac{(2x)^2}{(0.0482 - 2x)^2 \times (0.0933 - x)}$$

As K_c has small value, x should be low

$0.0482 - 2x \simeq 0.0482$

$0.0933 - x - x = 0.0933$

$$2 \times 10^{-37} = \frac{4x^2}{0.0482 \times 0.0933}$$

On solving, $x = 3.3 \times 10^{-21}$

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$[N_2O] = 2x = 6.6 \times 10^{-21}$

$[N_2] \simeq 0.0482\,M;\ [O_2] = 0.0993\,M$

42. One of the reaction that takes place in producing steel from iron ore is the reduction of iron(II) oxide by carbon monoxide to give iron metal and carbon dioxide, that is,

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g); \quad K_p = 0.265 \text{ atm}$$

at 1050 K. What are the equilibrium partial pressures of CO and CO_2, respectively, at 1050 K if the initial partial pressures are: $p_{CO} = 1.4$ atm and $p_{CO_2} = 0.80$ atm?

(A) p_{CO} = 1.739 atm and $p_{CO_2} = 0.461$ atm

(B) p_{CO} = 17.39 atm and $p_{CO_2} = 0.461$ atm

(C) p_{CO}= 1.79 atm and $p_{CO_2} = 0.46$ atm

(D) p_{CO} = 2.739 atm and $p_{CO_2} = 1.461$ atm

Solution

(A)

	$FeO(s)$	+	$CO(g)$	$\rightleftharpoons$	$Fe(s)$	+	$CO_2(g)$
$t = 0$			1.4 atm				0.8 atm
			$+x$				$-x$
At equilibrium			$(1.4 + x)$				$0.8 - x$

$$Q = \frac{0.8}{1.4} = \frac{4}{7} = 0.55$$

Since $Q > K$, system is not at equilibrium, and to attain equilibrium, it will shift in backward direction

$$K_p = 0.265 = \frac{0.8 - x}{1.4 + x}$$

On solving, we get x = 0.339

Therefore $[p_{CO}] = 1.739$ atm and $p_{CO_2} = 0.461$ atm

43. What are the values of ΔG^o and the equilibrium constant for the formation of NO_2 from NO and O_2 at 298 K?

$$NO(g) + \tfrac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$$

where

$\Delta G_f^o(NO_2) = 52.0\ \text{kJ mol}^{-1},\ \Delta G_f^o(NO) = 87.0\,\text{kJ mol}^{-1},\ \Delta G_f^o(O_2) = 0\,\text{kJ mol}^{-1}$.

(A) 35 kJ mol^{-1} and 1.365×10^6 (B) –35 kJ mol^{-1} and 1.365×10^{-6}

(C) –350 kJ mol^{-1} and 1.3×10^6 (D) –35 kJ mol^{-1} and 1.365×10^6

Solution

(D) $\Delta G^o = \Delta G^o_{NO_2} - \Delta G^o_{NG} = (52 - 87)\ \text{kJ mol}^{-1} = -35\,\text{kJ mol}^{-1}$

Since $\Delta G^o = -2.303\ RT\ \log K$

$$\log K = \frac{-35\,\text{kJ mol}^{-1}}{-8.301\,\text{kJ mol}^{-1} \times 298\,\text{K}}$$

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on solving $K = 1.365 \times 10^6$

44. A reversible chemical reaction is having two reactants in equilibrium. If the concentration of the reactants is doubled, then the equilibrium constant will
(A) also be doubled. (B) be halved.
(C) become one-fourth. (D) remain the same.
Solution
(D) K_c is a characteristic constant for the given reaction.

Multiple Correct Choice Type
1. Solid ammonium carbamate, $NH_4CO_2NH_2(s)$, dissociates into ammonia and carbon dioxide when it evaporates as shown by $NH_4CO_2NH_2(s) \rightleftharpoons 2NH_3(g) + CO_2(g)$. At 25°C, the total pressure of the gases in equilibrium with the solid is 0.116 atm. If 0.1 atm of CO_2 is introduced after equilibrium is reached, then
(A) final pressure of CO_2 will be less than 0.1 atm.
(B) final pressure of CO_2 will be more than 0.1 atm.
(C) pressure of NH_3 will decrease due to addition of CO_2.
(D) pressure of NH_3 will increase due to addition of CO_2.

Solution
(B, C) The pressure of NH_3 will decrease due to addition of CO_2, (backward, shifting Le Chatelier's principle). The pressure of CO_2 will be more than 0.1 atm.

2. The equilibrium of which of the following reaction will not be disturbed by the addition of an inert gas at constant volume?
(A) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ (B) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$
(C) $CO_2(g)+2H_2(g) \rightleftharpoons CH_3OH(g)$ (D) $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$
Solution
(A, B, C, D)

3. Consider the reactions given below. In which cases, will the reaction proceed toward the right by increasing the pressure?
(A) $4HCl(g) + O_2(g) \rightarrow 2Cl_2(g) + 2H_2O(g)$
(B) $Cl_2(g) + H_2O(g) \rightarrow 2HCl(g) + \frac{1}{2}O_2(g)$
(C) $CO_2(g) + 4H_2(g) \rightarrow CH_4(g) + 2H_2O(g)$
(D) $N_2(g) + O_2(g) \rightarrow 2NO(g)$
Solution
(A, C) With the increase in pressure, forward reaction is favored, because for forward reaction, lesser number of molecules are involved. This is in accordance with Le Chatelier's principle.

4. For the reaction $2A(g) + B(g) \rightleftharpoons 2C(g)$; $\Delta H = +13.6$ kJ, which of the following will increase the extent of the reaction at equilibrium
(A) increasing the temperature. (B) increasing the pressure.
(C) addition of catalyst. (D) removing C.
Solution
(B, C, D)

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5. The following reaction attains equilibrium at high temperature
$N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$
The yield of NO is affected by
(A) increasing the nitrogen concentration.
(B) decreasing the hydrogen concentration.
(C) compressing the reaction mixture.
(D) None of these.
Solution
(A, B, C). As the concentration of reaction (gaseous) is increased at equilibrium, reaction will go in the forward direction. Also, the above reaction is endothermic. Therefore, an increase in temperature will favor it.

6. If K_1 and K_2 are the equilibrium constants at temperature T_1 and T_2 and $T_2 > T_1$ then
(A) when $\Delta H = 0;\ K_2 = K_1$ (B) when $\Delta H > 0;\ K_2 > K_1$
(C) when $\Delta H < 0;\ K_2 < K_1$ (D) when $\Delta H > 0;\ K_2 < K_1$
Solution
(A, B, C) The relation is expressed as

$$\log\frac{K_2}{K_1} = \frac{\Delta H^{\circ}}{2.303R}\left[\frac{T_2 - T_1}{T_1T_2}\right]$$

If $T_2 > T_1$ and $\Delta H = 0$, we have $\log(K_2/K_1) = 0 \Rightarrow K_2/K_1 = 1 \Rightarrow K_2 = K_1$
If $T_2 > T_1$ and $\Delta H > 0$, we have $\log(K_2/K_1) > 0 \Rightarrow K_2/K_1 > 1 \Rightarrow K_2 > K_1$
If $T_2 > T_1$ and $\Delta H < 0$, we have $\log(K_2/K_1) < 0 \Rightarrow K_2/K_1 < 1 \Rightarrow K_2 < K_1$

7. $2CaSO_4(s) \rightleftharpoons 2CaO(s) + O_2(g)$, $\Delta H > 0$. Above equilibrium is established by taking sufficient amount of $CaSO_4(s)$ in a closed container at 1600 K. Then which of the following may be a correct option (Assume that solid $CaSO_4$ is present in the container in each case)?
(A) Moles of CaO(s) will increase with the increase in temperature.
(B) If the volume of the container is doubled at equilibrium, then partial pressure of $SO_2(g)$ will change at new equilibrium.
(C) If the volume of the container is halved, partial pressure of $O_2(g)$ at new equilibrium will remain same.
(D) If 2 mol of the He gas are added at constant pressure, then the moles of CaO(s) will increase.
Solution
(A, C, D)
(A) As the reaction is endothermic, therefore it will go in the forward direction. Hence, moles of CaO will increase.
(C) With the increase or decrease of volume, partial pressure of the gases will remain same.
(D) Due to the addition of inert gas at constant pressure, reaction will proceed in the direction in which more number of gaseous moles are formed.

8. In presence of a catalyst, what happens to the chemical equilibrium?

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(A) Energy of activation of the forward and backwards reactions is lowered by same amount.
(B) Equilibrium amount is not disturbed.
(C) Rates of forward and reverse reaction increase by the same factor.
(D) More product is formed.
Solution
(A, B, C) Catalyst only speeds up the reaction.

9. 138 g of $N_2O_4(g)$ is placed in 8.2 L container at 300 K. The equilibrium vapor density of mixture was found to be 30.67. Then (R = 0.082 L atm mol^{-1} K^{-1}).
(A) α = degree of dissociation of N_2O_4 = 0.25.
(B) K_p of $N_2O_4 \rightleftharpoons 2NO_2(g)$ will be 9 atm.
(C) Total pressure at equilibrium mixture will be 6.75 atm.
(D) The density of equilibrium mixture will be 16.83 g L^{-1}.
Solution
(B, C, D) The reaction is

	N_2O_4	$\rightleftharpoons$	$2NO_2$
$t = 0$	a		0
	$a(1-\alpha)$		$2a\alpha$

$$\text{Vapor density} = \frac{46}{1-\alpha} = 30.67 \Rightarrow \alpha = 0.5 = 50\%$$

$$\text{Total pressure} = = \frac{1.5 \times 1.5 \times 0.082 \times 300}{8.2} = 6.67\,\text{atm}$$

$$\text{So, } K_p = \frac{4\alpha^2}{1-\alpha^2}\,p = 9\,\text{atm}$$

$$\text{and density of mixture} = \frac{138}{8.2}\ \text{g L}^{-1} = 16.83\ \text{g L}^{-1}$$

10. $N_2F_4 \rightleftharpoons 2NF_2(g);\ \Delta H^o = 38.5\,\text{kJ}.$ Which of the following conditions will favor the formation of NF_2?
(A) Adding He to the equilibrium mixture at constant temperature and volume.
(B) Increasing the temperature.
(C) NF_2 gas is removed from the reaction mixture.
(D) Decreasing the pressure at constant temperature.
Solution
(B, C, D) In accordance with Le Chatelier's principle

11. Consider the following equilibria:
(I) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
(II) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
(III) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
Choose the correct statement(s):
(A) Addition of an inert gas at constant volume has no effect on all the three equilibria.
(B) Addition of an inert gas at constant pressure favors the forward reaction in (III), backward reaction in (I), and has no effect on (II).

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(C) Addition of an inert gas at constant pressure has no effect on equilibrium (II), but favors the forward reaction in (I) and backward reaction in (III).
(D) Addition of an inert gas has no effect on all the three equilibria at constant temperature and also at constant pressure.

Solution

(A, B)

12. The reactions in which the yield of the products cannot be increased by the application of high pressure are

(A) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(B) $NH_4HS\ (s) \rightleftharpoons NH_3(g) + H_2S(g)$

(C) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

(D) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Solution

(B, C)

13. For which of the following statements about the reaction quotient, *Q*, are correct?

(A) The reaction quotient, *Q*, and the equilibrium constant always have the same numerical value.

(B) *Q* may be > or < = K_c

(C) *Q* (numerical value) varies as reaction proceeds.

(D) *Q* = 1 at equilibrium

Solution

(B, C)

14. Consider the exothermic reaction

$$CoCl_4^{2-}(aq) + 6H_2O(l) \rightleftharpoons [Co(H_2O)_6]^{2+}(aq) + 4Cl^-(aq)$$

[Consider initially only $CoCl_4^{2-}$ is added to water]

(a) If HCl is added to the above equilibrium, then concentration of $CoCl_4^{2-}(aq)$ will increase.

(b) On dilution with water, the concentration of $CoCl_4^{2-}(aq)$ will decrease.

(c) On doubling the volume of solution by addition of water, the final concentration of $[Co(H_2O)_6]^{2+}$ will be more than its initial concentration.

(d) On increasing temperature, the concentration of $CoCl_4^{2-}$ will decrease.

Solution

(A, B, D) On adding Cl^- ion, the equilibrium will shift in the backward direction. So, $[CoCl_4]^{2-}$ will increase on dilution and hence the volume will increase.

Now, $$K_c = \frac{n_{[CO(H_2O)_6]^{2+}} . n_{Cl^-}^4}{n_{CoCl^{2-}}} . \frac{1}{V^4}$$

So, *Q* will get decreased and equilibrium will shift in the forward direction. Although the amount of $[Co(H_2O)]^{2+}$ will get increased, but its concentration cannot get doubled as volume has increased. As the reaction is exothermic in nature, so increasing temperature will shift the equilibrium in backward direction.

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15. For the equilibrium $N_2O_4 \rightleftharpoons 2NO_2$, the value of equilibrium constant is 2×10^{-4} at a particular temperature if equilibrium is established by taking 2 mol of N_2O_4 in 1 L container then

(A) concentration of NO_2 at equilibrium is 0.02 M approximately.
(B) concentration of NO_2 at equilibrium is 0.01 M approximately.
(C) concentration of N_2O_4 at equilibrium is 2 M approximately.
(D) concentration of N_2O_4 at equilibrium is 1 M approximately.

Solution

(B, C) $N_2O_4 \rightleftharpoons 2NO_2$

$2 - x \qquad 2x$

$$K_p = \frac{4x^2}{2} = 2 \times 10^{-4} \Rightarrow x = 0.01$$

[NO] = $2x$ = 0.02 M, $[N_2O_4]$ = 2/1 = 2 M

16. Which of the following relationship(s) describe the quantitative effect of temperature on the equilibrium constant?

(A) $\frac{d \ln K_p}{dT} = \frac{\Delta H}{RT^2}$ (B) $\frac{d \ln K_p}{dT} = \frac{\Delta E}{RT^2}$

(C) $\frac{d \ln K_c}{dT} = \frac{\Delta E}{RT^2}$ (D) $\frac{d \ln K_c}{dT} = \frac{\Delta H}{RT^2}$

Solution

(A, C)

Assertion–Reasoning Type

Choose the correct option from the following:

(A) Statement 1 is True, Statement 2 is True; Statement 2 is a correct explanation for Statement1.
(B) Statement 1 is True, Statement 2 is True; Statement 2 is NOT a correct explanation for Statement1
(C) Statement 1 is True, Statement 2 is False.
(D) Statement 1 is False, Statement 2 is True.

1. **Statement 1:** Addition of an inert gas at constant pressure to dissociation equilibrium of $PCl_5 \rightleftharpoons PCl_3(g) + Cl_2(g)$ favors forward reaction.

Statement 2: $K_C = \frac{\alpha^2 V}{1-\alpha}$ for dissociation equilibrium of PCl_5, where α is degree of dissociation of PCl_5.

Solution

(C) $K_p = \frac{n_{PCl_3} \times n_{Cl_2}}{n_{PCl_5}} \times \left(\frac{p}{n_T}\right)^1$ with addition of more and more inert gas n_T will increase that leads to decrease in the value of p/n_T, and hence to keep K_p constant, moles of the product would increase, that is, α increases.

2. **Statement 1:** In the Haber process, $N_2 + 3H_2 \rightleftharpoons 2NH_3$; $\Delta H < 0$. Then concentration of NH_3 increases by decreasing the temperature and increasing the pressure.

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Statement 2: For an exothermic reaction, equilibrium constant increases by increasing the temperature.
Solution
(C) In the exothermic reaction, equilibrium constant increases by decreasing the temperature.

3. Statement 1: The equilibrium amount of NO will be less with an increase in temperature.
Statement 2: Heat is absorbed by increasing the temperature and the equilibrium shifts toward left.
Solution
(A)

4. Statement 1: $\Delta G = \Delta G^o + 2.303RT \log_{10}Q$, where Q is reaction quotient.
Statement 2: Q may be greater or lesser than K or equal to K if $\Delta G = 0$.
Solution
(B) $Q < K_c$ forward reaction; $Q = K_c$ equilibrium; $Q > K_c$ backward direction.

5. Statement 1: Pure liquids and solids can be ignored while writing the equilibrium constant expression.
Statement 2: The concentration of pure liquids and solids does not remain constant.
Solution
(C) The concentration of pure liquids and solids remain constant, and these terms are amalgamated (or merged) in the equilibrium constant. That is why pure liquids and solids are ignored while writing the equilibrium constant expressions.
In the first step of Ostwald's process for synthesis of nitric acid, ammonium is oxidized to NO by the reaction

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H = O(g); \Delta H° = -905.6 \text{ kJ}$$

6. Statement 1: The equilibrium constant for a reaction having positive ΔH^o increases with an increase of temperature.
Statement 2: The temperature dependence of the equilibrium constant is related to ΔH^o and not ΔS^o for the reaction.
Solution
(B) $\log_{10} = \frac{k_2}{k_1} = \frac{\Delta H}{2.303R}\left[\frac{T_2 - T_1}{T_1T_2}\right]$

7. Statement 1: The equilibrium attained in a closed vessel remains unaltered by the addition of $CaCO_3(s)$, which is as follows: $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2\uparrow$
Statement 2: Heterogeneous reaction is unaffected by solid and liquid.
Solution
(A)

8. Statement 1: $SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g) + \text{heat}$. Forward reaction is favored at high temperature and low pressure.
Statement 2: Reaction is exothermic.
Solution

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(D) Reaction is exothermic, and low temperature favors forward reaction. High pressure favors forward reaction; it is accompanied by a decrease in the number of moles.

Comprehension Type

Read the paragraphs and answer the questions that follow.

Paragraph I

The difference in free energy of the reaction when all the reactants and products are in STP (1 atm and 298 K) is ΔG^o, and K_c and K_p are the thermodynamic equilibrium constants of the reaction. The relationship between them is as follows:

$$\Delta G^o = -2.303RT\log K_C$$

and $$\Delta G^o = -2.303RT \log K_p \text{ (in case of ideal gas)}$$

This equation relates the equilibrium constant of a reaction to a thermodynamic property. It is sometimes easier to calculate the free energy in a reaction rather than to measure the equilibrium constant. The standard free energy change can be calculated as

$$\Delta G^o = \Delta H^o - T\Delta S^o$$

where ΔH^o is the standard enthalpy change and ΔS^o is the standard entropy change.

1. Which of the following statement is correct for a reversible process in state of equilibrium?

(A) $\Delta G = -2.303RT \log K$ (B) $\Delta G = 2.303RT \log K$

(C) $\Delta G^o = -2.303RT \log K$ (D) $\Delta G^o = 2.303RT \log K$

Solution

(C)

2. Calculate the equilibrium concentration ratio of C to A if 2.0 mol each of A and B were allowed to come to equilibrium at 300 K

$$A + B \rightleftharpoons C + D;\ \Delta G^o = 460 \text{ cal}$$

(A) 1.0 (B) 0.5 (C) 0.8 (D) 0.679

Solution

(D) $\Delta G^o = -2.303RT \log k$

$\log k = \bar{1}.665 \Rightarrow k$ = Antilog of $\bar{1}.665 = 0.4625$

	A	+	B ⇔	C	+	D
Initial concentration	2		2	0		0
At equilibrium	$2-x$		$2-x$	x		x

$$K_e = \frac{x^2}{(2-x)^2} = 0.4625 \Rightarrow \frac{[C]}{[A]} = \frac{x}{2-x} = 0.68$$

3. At 490°C, the value of equilibrium constant K_p is 45.9 in the reaction

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

Calculate the value of ΔG^o for the reaction at that temperature

(A) –3.5 kcal (B) 3.5 kcal (C) 5.79 kcal (D) –5.79 kcal

Solution

(D) $\Delta G^o = -2.303RT \log K_p = -5.79$ kcal

Paragraph II

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Chemical equilibrium is a dynamic state in which the composition of a system does not change, and net reaction is zero. In reversible reactions, a stage is reached when the rate of transformation of reactants into products equals the rate of transformation of products into reactants. At this stage, the composition of reactants and products does not change with time. The reaction has not stopped, but the rates of forward and reverse reactions are equal. For a general reaction A(g) + B(g) $\rightleftharpoons$ C(g) + D(g), we have

$$K_C = \frac{[C][D]}{[A][B]} \quad \text{and} \quad K_p = \frac{p_C \times p_D}{p_A \times p_B}$$

where K_C is the equilibrium constant for the ratio of the concentrations of the product to reactants and K_p is the equilibrium constant for the ratio of partial pressure and of products to reactants. The relation between K_p and K_c is as follows.

$$K_p = K_C (RT)^{\Delta n}$$

4. The mass ratio of steam and hydrogen is found to be 1:2 at equilibrium.

$$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$$

The value of equilibrium constant (K_c) for the above reaction is

(A) 3.05×10^3 (B) 1.05×10^5 (C) 0.75×10^2 (D) 2.42×10^7

Solution

(B) As $\Delta n = 0$, so $K_c = K_p = K_x$

$$K_x = \frac{x_{H_2}^4}{x_{H_2O}^4} = \frac{(2/2)^4}{(1/18)^4} = 18^4 = 1.05 \times 10^5$$

5. For the reaction $NH_2COONH_4(s) \rightleftharpoons 2NH_3(g) + CO_2(g)$, the equilibrium constant $K_p = 2.9 \times 10^{-5}$ atm^3. The total pressure of gases at equilibrium when 1 mol of reactant was heated will be

(A) 0.0194 atm (B) 0.0388 atm (C) 0.0580 atm (D) 0.0667 atm

Solution

(C) $K_p = (p_{NH_3})^2 \times p_{CO_2} = (2p)^2 \times p = 4p^3$

$$p = \left(\frac{K_p}{4}\right)^{1/3} = 0.0667 \text{ atm}$$

6. K_c for the reaction $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) + \frac{1}{2}Br_2(g) \rightleftharpoons NOBr(g)$ from the following information at 298 K

$$2NO(g) \rightleftharpoons N_2(g) + O_2(g) \quad K_1 = 2.4 \times 10^{30}$$

$$NO(g) + \tfrac{1}{2}Br_2(g) \rightleftharpoons NOBr(g); \quad K_2 = 1.4$$

(A) 3.15×10^{-9} (B) 6.35×10^{-18}

(C) 9.03×10^{-16} (D) 17×10^{-17}

Solution

(C) $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g) \quad K_C = \left(\frac{1}{2.4 \times 10^{30}}\right)^{1/2}$

$$NO(g) + \tfrac{1}{2}Br_2(g) \rightleftharpoons NOBr(g) \qquad K_2 = 1.4$$

$$\tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) + \tfrac{1}{2}Br_2(g) \rightleftharpoons NOBr(g) \quad K = K_c \times K_2 = 9.03 \times 10^{-16}$$

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7. The equilibrium constant for the reaction $2SO_2 + O_2 \rightleftharpoons 2SO_3$ at 1000 K is 3.5. What would the partial pressure of oxygen gas to give equal moles of SO_2 and SO_3?
(A) 0.29 atm (B) 3.5 atm (C) 0.53 atm (D) 1.87 atm
Solution
(A) $K_p = \dfrac{p_{SO_3}^2}{p_{SO_2}^2 . p_{O_2}}$

Therefore $p_{O_2} = \dfrac{1}{K_p} = \dfrac{1}{3.5}\text{atm} = 0.29 \text{ atm}$

Paragraph III
In dissociation reactions involving gaseous molecules, the total mass remains same and the number of moles increases as a result of the reaction. Hence, the average molecular weight and vapor density decrease. So, the volume increases at constant temperature and pressure, and hence the vapor density decreases. The relationship between initial vapor density and that at equilibrium can be found as follows:
At constant temperature and constant pressure for a fixed mass of gaseous mixtures,

$$\text{Vapor density} \propto \rho \propto \frac{1}{V}$$

Now, for ideal gas, $pV = nRT$ or volume $\propto$ number of moles. So,

$$\text{Vapor density} \propto \frac{1}{\text{Number of moles}} \Rightarrow \frac{D}{D_0} = \frac{n}{n_0}$$

where D_0 and D are the vapor densities in the beginning and at equilibrium, respectively, and n_0 and n are the number of moles in the beginning and at equilibrium, respectively. The vapor densities of PCl_5 at 200°C and 252°C are 70.2 and 57.2, respectively, at 1 atm.

8. From the given data, it can be interpreted that
(A) if degree of dissociation is more, then vapor density is more for PCl_5.
(B) more is the degree of dissociation, less is the vapor density of PCl_5.
(C) degree of dissociation and vapor density cannot be correlated.
(D) α is independent of temperature.
Solution
(B) More is the degree of dissociation, less is the vapor density of $PC1_5$

9. X, Y, and Z react in the 1:1:1 stoichiometric ratio. The concentration of X, Y, and Z was found to vary with time as shown in Fig. 8.6.

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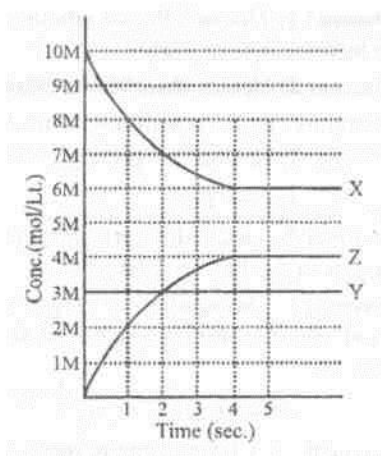

Figure 8.6

Which of the following equilibrium reaction represents the correct variation of concentration with time?

(A) $X(g) + Y(g) \rightleftharpoons Z(g)$ (B) $X(g) + Y(s) \rightleftharpoons Z(g)$

(C) $Z(g) + Y(s) \rightleftharpoons X(g)$ (D) $Z(g) + X(g) \rightleftharpoons Y(g)$

Solution

(B) Clearly, concentration of Y is not changing with the time. Hence, it will be pure solid or liquid. Concentration of X is decreasing. Hence, it will be in reactant and Z will be in product.

10. Value of the equilibrium constant (K_c) for the equilibrium represented in above sketch will be

(A) $\frac{9}{2}$ (B) $\frac{21}{4}$ (C) $\frac{2}{3}$ (D) $\frac{12}{7}$

Solution

(C) $K_C = \frac{[Z(g)]_{eq}}{[X(g)]_{eq}} = \frac{4}{6} = \frac{2}{3}$

11. If above equilibrium is established in a 2 L container by taking reactants in sufficient amount, then how many moles of component Y must have reacted for establishment of the equilibrium.

(A) 0 (B) 6 (C) 12 (D) 8

Solution

(D) From the graph, we can see that 4 mol L^{-1} of Z is formed. Therefore, 4×2 mol of solid will be required.

Integer Answer Type

The answer is a **non-negative integer**.

1. For the reaction: $CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$, K is 0.63 at 700°C and 1.66 at 1000°C. The average ΔH^o for the temperature range considered is ___.

Solution

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(8) Given that $T_1 = 700 + 273 = 973\ \text{K}$, $T_2 = 1000 + 273 = 1273\ \text{K}$, $K_1 = 0.63,\ K_2 = 1.66$

Using the van't Hoff equation

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

$$\log\left(\frac{1.66}{0.63}\right) = \frac{\Delta H^\circ}{2.303 \times 1.99}\left(\frac{1273 - 973}{1273 \times 973}\right) \Rightarrow \Delta H^\circ = 8.0 \times 10^3\ \text{cal} = 8.0\ \text{kcal}$$

Note: The units of R and ΔH must be same.

2. Equilibrium constant for the reaction $A_3(g) + 3B_2(g) \rightleftharpoons 3AB_2(g)$ is 64.0. Then the equilibrium constant for the reaction $\frac{1}{3}A_3(g) + B_2(g) \rightleftharpoons AB_2(g)$ will be ____.

Solution

(4) $K_C' = (64)^{1/3} = 4$

3. When NO and NO_2 are mixed, the following equilibria are readily obtained:

$$2NO_2 \rightleftharpoons N_2O_4;\ K_p = 6.8\ \text{atm}^{-1};\ \text{and}\ NO + NO_2 \rightleftharpoons N_2O_3$$

In an experiment when NO and NO_2 are mixed in the ratio of 1:2, then the total pressure was 5.05 atm and the partial pressure of N_2O_4 was 1.7 atm. Calculate the equilibrium partial pressure of NO in atmospheres.

Solution

(2) For 1 equilibrium $2NO_2 \rightleftharpoons N_2O_4$

$$K_p = \frac{p'_{N_2O_4}}{(p_{NO_2})^2} = 6.8 \qquad (1)$$

$p'_{N_2O_4} = 1.7\ \text{atm}$ therefore by Eq. (1) $p'_{NO_2} = 0.5\ \text{atm}$.

The equilibria are maintained using NO and NO_2 in the ratio of 1:2

For 2 equilibria

	NO	+	NO_2	$\rightleftharpoons$	N_2O_3
Initial pressures	p		$2p$		
Pressures at equilibria	$(p - x)$		$(2p - x - 3.4)$		x

3.4 atm of NO_2 are used for 1 equilibrium to have $p'_{N_2O_4} = 1.7\ \text{atm}$

At equilibrium	$(p - x)$	0.5	x

(Since p'_{NO_2} is same for both the equilibria, as both reactions are at equilibrium at a time)

Total pressure at equilibrium $= p'_{NO} + p'_{NO_2} + p'_{N_2O_3} + p'_{N_2O_4} = p - x + 0.5 + x + 1.7$

$$5.05 = p + 2.20 \Rightarrow p = 2.85\ \text{atm}$$

$$2p - x - 3.4 = 0.5 \Rightarrow 2 \times 2.85 - x - 3.4 = 0.5 \Rightarrow x = 1.80\ \text{atm} \approx 2\ \text{atm}$$

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4. Ammonium carbamate when heated to 200°C gives a mixture of vapors ($NH_2COONH_4 \rightleftharpoons 2NH_3 + CO_2$) with a density 13.0. The degree of dissociation of ammonium carbamate is ___.

Solution

(1) Density of ammonium carbamate before dissociation

$$D = \frac{\text{Molecular weight}}{2} = \frac{78}{2} = 39$$

Density after dissociation (d) = 13.0 (given). Hence,

$$\alpha = \frac{D-d}{(n-1)d} = \frac{39-13}{(3-1)\times 13} = \frac{26}{26} = 1$$

So, ammonium carbamate will be 100% dissociated.

5. For the equillibrium $N_2O_4 \rightleftharpoons 2NO_2$,

$$(\Delta G^{o}_{N_2O_4})_{298\,K} = 100\ \text{kJ mol}^{-1} \text{ and } (\Delta G^{o}_{NO_2})_{298\,K} = 50\ \text{kJ mol}^{-1}$$

When 5 mol L^{-1} of each in taken, calculate the value of ΔG for the reaction at 298 K.

Solution

(4)

	N_2O_4	$\rightleftharpoons$	$2NO_2$;	$\Delta G^{o}_{N_2O_4} = 100\,\text{kJ mol}^{-1}$
Concentration at $t = 0$	5		5	$\Delta G^{o}_{NO_2} = 50\ \text{kJ mol}^{-1}$

ΔG^{o} for reaction = $2 \times \Delta G^{o}_{NO_2} - \Delta G^{o}_{N_2O_4} = 2\times 50 - 100 = 0$

$$\Delta G = \Delta G^{o} + 2.303RT \log Q$$

Now, $\Delta G = 0 + 2.303 \times 8.314 \times 10^{-3} \times 298 \log \frac{5^5}{5} = +3.99\ \text{kJ} \approx 4\ \text{kJ}$

6. If concentrations of SO_2 and O_2 in the equilibrium reaction, $2SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$ are quadrupled, the concentration of SO_3 now will be ___ times.

Solution

(8) For the reaction

$$K_C = \frac{[SO_3]^2_{\text{initial}}}{[SO_2]^2[O_2]} \quad (1)$$

When concentrations of SO_2 and O_2 are quadrupled, the expression becomes

$$K_C = \frac{[SO_3]^2_{\text{final}}}{[4SO_2]^2[4O_2]} \quad (2)$$

Dividing Eq. (2) by Eq. (1), we get

$$\frac{[SO_3]^2_{\text{final}}}{[4SO_2]^2[4O_2]} = \frac{[SO_3]^2_{\text{initial}}}{[SO_2]^2[O_2]} \Rightarrow \frac{[SO_3]^2_{\text{final}}}{[SO_3]^2_{\text{initial}}} = \frac{[4SO_2]^2[4O_2]}{[SO_2]^2[O_2]} = 16 \times 4 = 64$$

Hence, $$\frac{[SO_3]_{\text{final}}}{[SO_3]_{\text{initial}}} = 8$$

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7. For the reaction involving oxidation of ammonia by oxygen to form nitric oxide and water vapor, the equilibrium constant has the units $(\text{bar})^n$. Then n is ____.

Solution

(1) $4NH_3(g) + 7O_2(g) \rightleftharpoons 4NO_2(g) + 6H_2O(g)$

Water and oxygen will not contribute to partial pressure.

$$Kp = \frac{(pNO_2)^4}{(pNH_3)^4} = 1$$

Answer should be -1.

Matrix–Match Type

1. Match the reactions of the column I with the relations in Column II

Column I	**Column II**
(A) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	(p) $K_p = K_c(RT)$
(B) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	(q) $K_p = K_c(RT)^2$
(C) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$	(r) $K_p = K_c(RT)^{-2}$
(D) $NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$	(s) $K_p = K_c$
	(t) $K_p = K_c(RT)^{-1}$

Solution

(A) → (s); (B) → (r); (C) → (p); (D) → (q)

Using the relation $K_p = K_c(RT)^{\Delta n}$

2.

Match the relation with the prediction of equilibrium.

Column I	**Column II**
(A) $Q = K$	(p) Reaction is nearer to completion
(B) $Q < K$	(q) Reaction is not at equilibrium
(C) $Q > K$	(r) Reaction is fast in forward direction;
(D) $K \gg 1$	(s) Reaction at equilibrium
	(t) Rate of forward reaction = rate of backward reaction

Solution

(A) → (s, t); (B) → (q, r); (C) → (q); (D) → (p)

3. Match the chemical reaction in equilibrium with the pressure dependent on degree of dissociation of the reaction:

Column I	**Column II**
(A) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	(p) $\alpha \propto \sqrt{p}$
(B) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$	(q) $\alpha \propto \sqrt{1/p}$
(C) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$	(r) $\alpha \propto p$
(D) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$	(s) $\alpha \propto p^0$
	(t) $K_p = K_c$

Solution

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(A) → (r); (B) → (p); (C) → (q); (D) → (s, t)

4. Match the effect of change administered to its effect on equilibrium.

Column I	**Column II**
(A) For the equilibrium $NH_4I(s) \rightleftharpoons NH_3(g) + HI(g)$, if pressure is increased at equilibrium	(p) Forward shift
(B) For the equilibrium $N_2 + 3H_2 \rightleftharpoons 2NH_3$ At equilibrium volume is increases at equilibrium	(q) No change
(C) For the equilibrium $H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g)$ inert gas is added at constant pressure at equilibrium	(r) Backward shift
(D) For the equilibrium $PCl_5 \rightleftharpoons PCl_3 + Cl_2$, Cl_2 is removed at equilibrium	(s) Final pressure is more than initial pressure

Solution

(A) → (r),(s); (B) →(r); (C) → (q); (D) → (p),(s)

(A) As $\Delta n > 0$ therefore if $p \uparrow$, reaction will go in the backward direction.

(B) As $\Delta n < 0$ therefore if $V \uparrow$, $p \downarrow$ reaction will go in the direction in which more number of gaseous moles are formed that is backward direction.

(C) As $\Delta n = 0$, hence no effect.

(D) If concentration of the product is decreased, reaction will go in the forward direction.

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5. Match the quantity with the effect/relation.

Column I	**Column II**
(A) $K_C < 1$	(p) Affected by temperature
(B) Degree of dissociation	(q) Affected by pressure
(C) Equilibrium constant	(r) $K_f < K_b$
(D) Melting of ice	(s) $K_f > K_b$

Solution

(A) → (r); (B) → (p, q); (C) → (p); (D) → (p, q)

(A) Backward reaction is favored.

(B) α is affected by both temperature and pressure.

(C) K_C is affected by temperature only.

(D) On increasing temperature and pressure, ice melts.

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