Question 11. A sample of HI (g) is placed in a flask at a pressure of 0.2 atm. At equilibrium partial pressure of HI (g) is 0.04 atm. What is K_p for the given equilibrium?

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

Answer:

$$pHI = 0.04\ atm,\ pH_2 = 0.08\ atm;\ pI_2 = 0.08\ atm$$

$$K_p = \frac{pH_2 \times pI_2}{p^2_{HI}} = \frac{(0.08\ atm) \times (0.08\ atm)}{(0.04\ atm) \times (0.04\ atm)} = \mathbf{4.0}$$

Question 12. A mixture of 1.57 mol of N_2, 1.92 mol of H_2 and 8.13 mol of NH_3 is introduced into a 20 L reaction vessel at 500 K. At this temperature, the equilibrium constant K_c for the reaction

$$N_2\ (g) + 3H_2\ (g) \rightleftharpoons 2NH_3\ (g)\ \text{is}\ 1.7 \times 10^{-2}.$$

Is this reaction at equilibrium? If not, what is the direction of net reaction?

Answer:

The reaction is: $N_2(g) + 3\ H_2(g) \rightleftharpoons 2NH_3(g)$

$$\text{Concentration quotient } (Q_c) = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(8.13/20\ mol\ L^{-1})^2}{(1.57\ /\ 20\ mol\ L^{-1}) \times (1.92\ /\ 20\ mol\ L^{-1})^3}$$

$$= \mathbf{2.38 \times 10^3}$$

The equilibrium constant (K_c) for the reaction = 1.7×10^{-2}

As $Q_c \neq K_c$; this means that the reaction is not in a state of equilibrium.

Question 13. The equilibrium constant expression for a gas reaction is,

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

Write the balanced chemical equation corresponding to this expression.

Answer: Balanced chemical equation for the reaction is 4

$$4\ NO\ (g) + 6H_2O(g) \rightleftharpoons 4\ NH_3\ (g) + 5\ O_2(g)$$

Question 14. If l mole of H_20 and 1 mole of CO are taken in a 10 litre vessel and heated to 725 K, at equilibrium point 40 percent of water (by mass) reacts with carbon monoxide according to equation.

$$H_2O + CO(g) \rightleftharpoons H_2\ (g) + CO_2(g)$$

Calculate the equilibrium constant for the reaction.

Answer:

Number of moles of water originally present = 1 mol

Percentage of water reacted =40%

Number of moles of water reacted = 1 x 40/100 = 0.4 mol
Number of moles of water left = (1 - 0.4) = 0.6 mole
According to the equation, 0.4 mole of water will react with 0.4 mole of carbon monoxide to form 0.4 mole of hydrogen and 0.4 mole of carbon dioxide.
Thus, the molar cone, per litre of the reactants and products before the reaction and at the equilibrium point are as follows:

	$H_2O(g)$ +	$CO(g)$ $\rightleftharpoons$	$H_2(g)$ +	$CO_2(g)$
Initial moles/litre	$\frac{1}{10}$	$\frac{1}{10}$	0	0
Mole/litre at the equilibrium point	$\frac{1-0.4}{10}=\frac{0.6}{10}$	$\frac{1-0.4}{10}=\frac{0.6}{10}$	$\frac{0.4}{10}$	$\frac{0.4}{10}$

Applying law of chemical equilibrium,

$$\text{Equilibrium constant } (K_c) = \frac{[H_2(g)][CO_2(g)]}{[H_2O(g)][CO(g)]} = \frac{\left(\frac{0.4}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.4}{10}\text{mol L}^{-1}\right)}{\left(\frac{0.6}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.6}{10}\text{mol L}^{-1}\right)}$$

$$= \frac{0.16}{0.36} = \mathbf{0.44}$$

Question 15. What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?

$$2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)\ ;\ K_c = 0.14$$

Answer:
Suppose at equilibrium, the molar concentration of both I_2 (g) and Cl_2 (g) is x mol L^{-1}.

	$2ICl(g)$ $\rightleftharpoons$	$I_2(g)$ +	$Cl_2(g)$
Initial molar conc.	0.78 M	0	0
Eqn. molar conc.	$(0.78 - 2x)$M	x	x

$$K_c = \frac{[I_2(g)][Cl_2(g)]}{[ICl(g)^2]} = \frac{(x)\times(x)}{(0.78-2x)^2}$$

$$\frac{x}{(0.78-2x)} = (0.14)^{1/2} = 0.374 \quad \text{or} \quad x = 0.374\,(0.78 - 2x)$$

$$x = 0.292 - 0.748x \quad \text{or} \quad 1.748x = 0.292\ ;\ x = \frac{0.292}{1.748} = 0.167$$

$$[ICl] = (0.78 - 2 \times 0.167) = (0.78 - 0.334) = \mathbf{0.446\ M}$$
$$[I_2] = 0.167\text{ M}\ ;\ [Cl_2] = \mathbf{0.167\ M}$$

Question 16. K =0.04 atm at 898 K for the equilibrium shown below. What is the equilibrium concentration ok C_2H_6 when it is placed in a flask at 4 atm pressure,and allowed to come to equilibrium.

$$C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$$

Answer:
The equilibrium in the reaction is:

	$C_2H_6(g)$ $\rightleftharpoons$	$C_2H_4(g)$ +	$H_2(g)$
Initial pressure:	4 atm	0	0
Eqn. pressure:	$(4-p)$ atm	p atm	p atm

$$K_p = \frac{p_{C_2H_4}\times p_{H_2}}{p_{C_2H_6}} \quad \text{or} \quad 0.04 = \frac{p^2}{(4-p)}$$

$$p^2 = 0.04\,(4-p) \quad \text{or} \quad p^2 + 0.04\,p - 0.16 = 0$$

$$p = \frac{(-0.04)\pm\sqrt{0.0016-4\,(-0.16)}}{2}$$

$$= \frac{(-0.04)\pm 0.8}{2} = \frac{0.76}{2} = 0.38$$

Equilibrium pressure or concentration of C_2H_6 = (4 – 0.38) = 3.62 atm.

Question 17. The ester, ethyl acetate is formed by the reaction of ethanol and acetic acid and the equilibrium is represented as:
$CH_3COOH(l) + C_2H_5OH(l) \rightarrow CH_3COOC_2H_5(l) + H_2O(l)$
(i) Write the concentration ratio (concentration quotient) Q for this

reaction. Note that water is not in excess and is not a solvent in this reaction.
(ii) At 293 K, if one starts with 1.000 mol of acetic acid and 0.180 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.
(iii) Starting mth 0.50 mol of ethanol and 1.0 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?
Answer:

(i) The concentration ratio (Concentration quotient) Q_c for the reaction is:

$$Q_c = \frac{[CH_3COOC_2H_5(l)][H_2O(l)]}{[CH_3COOH(l)][C_2H_5OH(l)]}$$

(ii)

	$CH_3COOH(l)$ +	$C_2H_5OH(l)$ ⇌	$CH_3COOC_2H_5(l)$ +	$H_2O(l)$
Initial molar conc.	1.0 mol	0.18 mol	0	0
Molar conc. at equilibrium point	(1 – 0.171) = 0.829 mol	(0.18 – 0.171) = 0.009 mol	0.171 mol	0.171 mol

Applying Law of Chemical equilibrium,

$$K_c = \frac{[CH_3COOC_2H_5](l)[H_2O(l)]}{[CH_3COOH(l)][C_2H_5OH(l)]}$$

$$= \frac{(0.171\ \text{mol}) \times (0.171\ \text{mol})}{(0.829\ \text{mol})\ (0.009\ \text{mol})} = \mathbf{3.92}$$

(iii)

	$CH_3COOH(l)$ +	$C_2H_5OH(l)$ ⇌	$CH_3COOC_2H_5(l)$ +	$H_2O(l)$
Initial molar conc.	1.0 mol	0.5 mol	0.214 mol	0.214 mol
Molar conc. at equilibrium	1.0 – 0.214 = 0.786 mol	0.5 – 0.214 = 0.286 mol		

$$Q_c = \frac{[CH_3COOC_2H_5(l)][H_2O(l)]}{[CH_3COOH(l)][C_2H_5OH(l)]}$$

$$= \frac{(0.214\ \text{mol}) \times (0.214\ \text{mol})}{(0.786\ \text{mol})\ (0.286\ \text{mol})} = \mathbf{0.204}$$

Since Q_C is less than K_C this means that the equilibrium has not been reached. The reactants are still taking part in the reaction to form the products.

Question 18. A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was reached, the concentration of PCl_5 was found to be 0.5×10^{-1} mol L^{-1}. If K_C is 8.3×10^{-3} what are the concentrations of PCl_3 and Cl_2 at equilibrium?
Answer:
Let the initial molar concentration of PCl_5 per litre = x mol
Molar concentration of PCl_5 at equilibrium = 0.05 mol
∴ Moles of PCl_5 decomposed = (x - 0.05) mol
Moles of PCl_3 formed = (x - 0.05) mol
Moles of Cl_2 formed = (x - 0.05) mol
The molar conc./litre of reactants and products before the reaction and at the equilibrium point are:

	PCl_5 ⇌	PCl_3 +	Cl_2
Initial moles/litre	x	0	0
Moles/litre at eqm. point	0.05	$(x - 0.05)$	$(x - 0.05)$

Equilibrium constant (K_c) = 8.3×10^{-3} = 0.0083
Applying Law of chemical equilibrium,

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}; \quad 0.0083 = \frac{(x - 0.05) \times (x - 0.05)}{0.05}$$

$$(x - 0.05)^2 = 0.0083 \times 0.05 = 4.15 \times 10^{-4}$$
$$(x - 0.05) = (4.15 \times 10^{-4})^{1/2} = 2.037 \times 10^{-2} = 0.02 \text{ moles}$$
$$x = 0.05 + 0.02 = 0.07 \text{ mol}$$

The molar concentration per litre of PCl_3 at eqm. = 0.07 – 0.05 = 0.02 mol
The molar concentration per litre of Cl_2 at eqm. = 0.07 – 0.05 = 0.02 mol.

Question 19. One of the reactions that takes place in producing steel from iron ore is the reduction of iron
(II) oxide by carbon monoxide to give iron metal and CO_2
$FeO(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$; K_p = 0.265 atm at 1050 K
What are the equilibrium partial pressures of CO and CO_2 at 1050 K if the initial pressures are: P_{CO} = 1.4 atm and P_{CO2} = 0.80 atm?
Answer:

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$$

Initial pressure: 1.4 atm 0.8 atm

$$Q_p = \frac{p_{CO_2}}{p_{CO}} = \frac{(0.8 \text{ atm})}{(1.4 \text{ atm})} = 0.571$$

Since $Q_p > K_p (0.265)$, this means that the reaction will move in the backward direction to attain the equilibrium. Therefore, partial pressure of CO_2 will decrease while that of CO will increase so that the equilibrium may be attained again. Let p atm be the decrease in the partial pressure of CO_2. Therefore, the partial pressure of CO will increase by the same magnitude i.e., p atm.

$$p_{CO_2} = (0.8 - p) \text{ atm}; \; p_{CO}(g) = (1.4 + p) \text{ atm}$$

At equilibrium, $$K_p = \frac{p_{CO_2}}{p_{CO}} = \frac{(0.8-p)\text{ atm}}{(1.4+p)\text{ atm}} = \frac{(0.8-p)}{(1.4+p)}$$

or $$0.265 = \frac{(0.8-p)}{(1.4+p)}$$

$$0.371 + 0.265\,p = 0.8 - p \quad \text{or} \quad 1.265\,p = 0.8 - 0.371 = 0.429$$

$$p = 0.429/1.265 = 0.339 \text{ atm}$$

$$(p_{CO})_{eq} = (1.4 + 0.339) = \mathbf{1.739 \text{ atm}}$$

$$(p_{CO_2})_{eq} = (0.8 - 0.339) = \mathbf{1.461 \text{ atm}}$$

Question 20.

Equilibrium constant K_c for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500 K is 0.061. At particular time, the anaylsis shows that the composition of the reaction mixture is: 3.0 mol L^{-1} of N_2; 2.0 mol L^{-1} of H_2; 0.50 mol L^{-1} of NH_3. Is the reaction at equilibrium? If not, in which direction does the reaction tend to proceed to reach the equilibrium?

Answer:

The given reaction is:

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

According to available data.

$$N_2 = [3.0]\,;\; H_2 = [2.0]\,;\; NH_3 = [0.50]$$

$$Q_c = \frac{[NH_3(g)]^2}{[N_2(g)][H_2(g)]^3} = \frac{[0.50]^2}{[3.0][2.0]^3} = \frac{0.25}{24} = \mathbf{0.0104.}$$

Since the value of Q_c is less than that of K_c (0.061), the reaction is not in a state of equilibrium. It will proceed in the **forward direction** till Q_c becomes the same as K_c.

*********** END **********