23 April 2018 H.W Tuesday

# Chapter: 9

## "Chemical Equilibrium"

### Self Assessment Exercise: 9.3

Q. Determine the units of equilibrium constants for the following reactions.

1. $N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

$$= \frac{[mol \cdot dm^{-3}]^2}{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}$$

$K_c$ = no units

2. $H_{2(g)} + CO_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$

$$K_c = \frac{[H_2O][CO]}{[H_2][CO_2]}$$

$$= \frac{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}$$

$K_c$ = no units.

3. $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

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

$$= \frac{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}{[mol \cdot dm^{-3}]}$$

$K_c = [mol \cdot dm^{-3}]$.

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4. $CO_{(g)} + 2H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$$

$$k_c = \frac{[mol\ dm^{-3}]}{[mol\ dm^{-3}][mol \cdot dm^{-3}]^2}$$

$$k_c = \frac{1}{(mol \cdot dm^{-3})^2}$$

$$= mol^{-2} \cdot dm^6$$

$$= dm^6 \cdot mol^{-2}$$

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# Chapter: 9

# "Chemical Equilibrium"

## Review Exercise:

Q2: Give short answers:

(i) Differentiate between forward and reverse reactions.

| Forward Reaction | Reverse Reaction |
|---|---|
| **Direction of Reaction** | |
| Forward reaction always proceeds from left to right. | Reverse reaction always proceeds from right to left. |
| **Resultant Specie** | |
| In forward reaction reactants react to produce products. | In reverse reaction reaction products react to produce reactants. |
| **Rate of Reaction** | |
| Rate is fastest in the beginning and gradually slows down. | Its rate is zero in the beginning and gradually speeds up. |
| **Example** | |
| $H_2 + I_2 \rightleftharpoons 2HI$ | |
| In this reaction Hydrogen ($H_2$) and Iodine ($I_2$) react to produce hydrogen idide and produces product, it is forward reaction i.e: $H_2 + I_2 \rightarrow 2HI$ | In this reaction hydrogen iodide decompose to produce hydrogen and iodine and it moves from right to left. i.e: $2HI \rightarrow H_2 + I_2$. |

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(ii) What is chemical equilibrium?

## Chemical Equilibrium:

### Definition:

"A state of a chemical reaction in which forward and reverse reactions take place at the same rate is called chemical equilibrium."

### Explanation:

Chemical equilibrium is a dynamic equilibrium. This is because reactions do not stop when they come to equilibrium state. The individual molecules keep on reacting continuously. But there is no change in the actual amounts of reactants and products. This means concentration of reactants and products become constant at equilibrium stage.

(iii) State the law of Mass Action.

## Law of Mass Action:

### Statement:

"It states that the rate at which a substance reacts is directly proportional to its active mass and the rate at which the reaction proceeds

is directly proportional to the product of the active masses of the reactants."

**Explanation:**

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

According to law of mass action:

Rate of forward reaction $\propto [A]^a [B]^b$

$$R_f = k_f [A]^a [B]^b$$

Rate of reverse reaction $\propto [C]^c [D]^d$

$$R_r = k_r [C]^c [D]^d$$

(iv) **State conditions for equilibrium.**

**Conditions for Equilibrium:**

Equilibrium is reached whether pure reactants, pure products or a mixture of reactants and products is initially placed in a closed container. In any of these cases, a condition will eventually be reached in the container such that the forward and reverse reaction occur at exactly the same rate. This leads to situation where the concentration of reactants and products remain the same indefinitely, as long as the following conditions are observed:

**Concentration** of none of the reactant or product is changed.

**Temperature** of the system is kept constant.

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**Pressure or Volume** of the system is kept constant.

**Q3:** Following reaction can occur during lightning storms.

$$3O_{2(g)} \rightleftharpoons 2O_{3(g)}$$

For this reaction write:

(i) Equilibrium constant Expression.

According to the law of mass actions:

**Forward Reaction:**

Rate of Forward reaction $\propto [O_2]^3$

$$R_f = k_f [O_2]^3$$

**Reverse Reaction:**

Rate of reverse reaction $\propto [O_3]^2$

$$R_r = k_r [O_3]^2$$

**At Equilibrium state:**

Rate of forward reaction = Rate of reverse reaction

$$R_f = R_r$$

$$k_f [O_2]^3 = k_r [O_3]^2$$

$$\frac{k_f}{k_r} = \frac{[O_3]^2}{[O_2]^3}$$

This is the equilibrium expression.

(ii) Determine the units of equilibrium constant:

$$k_c = \frac{[O_3]^2}{[O_2]^3}$$

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$$K_c = \frac{[mol.dm^{-3}][mol.dm^{-3}]}{[mol.dm^{-3}][mol.dm^{-3}][mol.dm^{-3}]}$$

$$= \frac{1}{[mol.dm^{-3}]}$$

$$= [mol^{-1}\ dm^3]$$

$$K_c = dm^3.mol^{-1}.$$

(iii) Forward and reverse reactions.

Forward Reaction:

$$3O_{2(g)} \rightarrow 2O_{3(g)}$$

Reverse Reaction:

$$2O_{3(g)} \rightarrow 3O_{2(g)}$$

Q4: Coal reacts with hot steam to form CO and $H_2$. These substances react further in the presency of catalyst to give methane and water vapour.

$$CO_{(g)} + 3H_{2(g)} \rightleftharpoons CH_{4(g)} + H_2O_{(g)}$$

(i) Write forward and reverse reactions for it.

Forward Reaction:

$$CO_{(g)} + 3H_{2(g)} \rightarrow CH_{4(g)} + H_2O_{(g)}$$

Reverse Reaction:

$$CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{2(g)} + 3H_{2(g)}.$$

(ii) Derive $K_c$ expression for the reaction.

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According to law of mass action:

**Forward Reaction:**

Rate of forward reaction $\propto [CO][H_2]^3$

$$R_f = k_f [CO][H_2]^3$$

**Reverse Reaction:**

Rate of reverse reaction $\propto [CH_4][H_2O]$

$$R_r = k_r [CH_4][H_2O]$$

**In Equilibrium State:**

Rate of forward reaction = Rate of Reverse reaction

$$k_f [CO][H_2]^3 = k_r [CH_4][H_2O]$$

$$\frac{k_f}{k_r} = \frac{[CH_4][H_2O]}{[CO][H_2]^3} \Rightarrow K_c = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

This is required equilibrium expression. $\frac{k_f}{k_r} = K_c$

(iii) Determine units for $K_c$.

$$K_c = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

$$= \frac{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]^3}$$

$$= \frac{1}{[mol \cdot dm^{-3}]^2}$$

$$= [mol \cdot dm^{-3}]^{-2}$$

$$= [mol^{-2} \, dm^{6}]$$

$$= dm^6 \cdot mol^{-2}.$$

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# Chapter: 9
## "Chemical Equilibrium"

### Review Exercise:

Q5. Write equilibrium constant expression for each of following reactions.

(i) $H_2O_{(g)} \rightleftharpoons H_{2(g)} + \frac{1}{2} O_{2(g)}$

$$K_c = \frac{[H_2][O_2]^{1/2}}{[H_2O]}$$

(ii) $CO_{(g)} + 2H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$$

(iii) $COCl_{2(g)} \rightleftharpoons CO_{(g)} + Cl_{2(g)}$

$$K_c = \frac{[CO][Cl_2]}{[COCl_2]}$$

(iv) $4HCl_{(g)} + O_{2(g)} \rightleftharpoons 2Cl_{2(g)} + 2H_2O_{(g)}$

$$K_c = \frac{2[Cl_2]^2[H_2O]^2}{[HCl]^4[O_2]}$$

Q6. Determine the units of equilibrium constants for following reactions.

(i) $COCl_{2(g)} \rightleftharpoons CO_{(g)} + Cl_{2(g)}$

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$$K_c = \frac{[CO][Cl_2]}{[COCl_2]}$$

$$= \frac{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}{[mol \cdot dm^{-3}]}$$

$$= mol \cdot dm^{-3}.$$

(ii) $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

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

$$= \frac{[mol \cdot dm^{-3}]^2}{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}$$

$$= \text{no units.}$$

(iii) $2H_{2(g)} + O_{2(g)} \rightleftharpoons 2H_2O_{(g)}$

$$K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]}$$

$$= \frac{[mol \cdot dm^{-3}]^2}{[mol \cdot dm^{-3}]^2[mol \cdot dm^{-3}]}$$

$$= \frac{1}{[mol \cdot dm^{-3}]}$$

$$= mol^{-1} \cdot dm^3$$

$$= dm^3 \cdot mol^{-1}.$$

(iv) $N_{2(g)} + 2O_{2(g)} \rightleftharpoons 2NO_{2(g)}$

$$K_c = \frac{[NO_2]^2}{[N_2][O_2]^2}$$

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$$K_c = \frac{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]}{[mol \cdot dm^{-3}][mol \cdot dm^{-3}]^2}$$

$$= \frac{1}{[mol \cdot dm^{-3}]}$$

$$= mol^{-1} \cdot dm^3$$

$$= dm^3 \cdot mol^{-1}.$$

**Q7: State the ways that equilibrium can be recognized.**

**Ways to Recognize Equilibrium:**

Equilibrium state of chemical reaction can be determined by determining concentration of reactants and products at regular intervals. When constant concentration of products and reactants is observed, the reaction is equilibrium. This can be done by physical methods as well as chemical methods such as titration, spectroscopy etc.

"Think-Tank"

Q1. Bromine chloride ($BrCl$) decomposes to form chlorine and bromine. For this reaction write.

(i) Chemical Equation.

$$2BrCl \rightleftharpoons Cl_2 + Br_2$$

(ii) $K_c$ expression.

$$K_c = \frac{[Cl_2][Br_2]}{[BrCl]^2}.$$

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(iii) Units of $K_c$.

$$K_c = \frac{[Cl_2][Br_2]}{[BrCl]^2}$$

$$= \frac{[mol\,dm^{-3}][mol\,dm^{-3}]}{[mol\,dm^{-3}]^2}$$

= no units.

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Q2: $K_c$ expression for reaction is given below:

$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

Chose reactant and product and derive the units of $K_c$ for this reaction.

$$N_2O_4 \rightleftharpoons 2NO_2$$

(Reactant) (Product)

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Q3: For which of following reactions are both reactants and products likely to be found when the reaction appears to be complete. Justify

(i) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

(ii) $2HF_{(g)} \rightleftharpoons H_{2(g)} + F_{2(g)}$

Ans: $2HF_{(g)} \rightleftharpoons H_{2(g)} + F_{2(g)}$

As equilibrium established only in reversible reaction and when the reaction is in equilibrium it contains both products and reactants. So the

reaction in which hydrogen flouride decompose to give hydrogen and flourine, both reactants and products likely to be found when the reaction appears to complete.

Cobalt chloride forms pink crystals ($COCl_3 \cdot 6H_2O$). When they are heated water is evolved and they turn blue ($COCl_2$). Can you use cobalt chloride as test for water. Argue.

$$COCl_3 \cdot 6H_2O \rightleftharpoons COCl_3 + 6H_2O$$

Cobalt chloride is used as a test for water. To check if the clourless liquid is water, test with piece of cobalt chloride paper. Water turns blue paper pink.

$$\underset{\text{(blue)}}{COCl_3} \rightleftharpoons \underset{\text{(pink)}}{COCl_3} + 6H_2O$$

## "Review Exercise"

Q2(v) What is the importance of equilibrium constant for a chemical reaction?

### Direction of Reaction:

As $K_c = \dfrac{[Product]}{[Reactant]}$

The direction of chemical reaction at any particular time can be predicted by means of [Product]/[Reactant] ratio leads by to one of the following three possibilities:

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(a) The ratio is less than $K_c$. This implies that more of the product is required to attain equilibrium, therefore the reaction will proceed in the forward direction.

(b) The ratio is greater than $K_c$. This means that reverse reaction will occur to attain the equilibrium.

(c) When the ratio is equal to $K_c$, the reaction is at equilibrium.

## Extent of Reaction:

(a) If the equilibrium constant is very large, this indicates that the reaction is almost complete.

(b) If the value of $K_c$ is very small, it reflects that the reaction doesnot proceed appreciably in the forward direction.

(c) If the value of $K_c$ is moderate this shows very little forward reaction.

## The Effect of External Conditions on the position of Equilibrium:

When a system reaches equilibrium it will remain in the same state indefinetely, if the conditions donot change. However, equilibrium state of a system is disturbed, if external conditions are

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changed e.g change of pressure, temperature, concentration of reactant and product - will alter the position of equilibrium.
Whenever the equilibrium is disturbed by changes in the conditions, the system will always tend to restore equilibrium.

**Concentration:**

It can be used to determine the equilibrium concentration of equilibrium mixture knowing the initial concentration of reactants.

Good
Neat Work