Q: Differentiate between reversible and irreversible reaction.

Reversible reaction	Irreversible reaction
Definition	
The reaction in which products can react together to reform original reactants is called reversible reaction.	The reaction in which products don't react together to reform original reactants is called irreversible reaction.
Direction	
It proceeds in forward as well as reverse direction.	It proceed only in one direction.
Completion	
It never goes completion.	It typically proceeds to completion.
Chemical Equation	
In their chemical equation, reactants and products are separated by a double half-headed arrow ($\rightleftharpoons$).	In their chemical equations, reactants and products are separated by a single-headed arrow ($\longrightarrow$).
Equilibrium	
They can reach state of equilibrium.	They are unable to reach equilibrium state because they typically go to completion.
Examples	
1) $N_2 + 3H_2 \rightleftharpoons 2NH_3$	1) $Mg + 2HCl \longrightarrow MgCl_2 + H_2$
2) $2SO_2 + O_2 \rightleftharpoons 2SO_3$	2) $F_2 + 2NaCl \longrightarrow 2NaF + Cl_2$

Chemical Equilibrium:

Q: What type of molecules are present at equilibrium state?

ans: At equilibrium, a mixture of reactants and products is present which is regarded as equilibrium mixture. Thus, both reactants and products are present at equilibrium state.

~~Q: Describe equilibrium state~~

Which reaction is called forward reaction

ans: Forward reaction:

In a reversible reaction, the reaction of reactants to form products is called forward reaction.

Written from:

It is written from ~~right~~ left to right.

Rate:

Its rate is fastest in the beginning (due to maximum concentration of reactants) and then gradually slows down.

Which reaction is called reverse reaction?

ans: Reverse reaction:

In a reversible reaction, the reaction of products to re-form original reactants is called reverse reaction.

Written from:

It is written from right to left.

Rate:

Its rate is ~~fastest~~ zero in the beginning (due to zero concentration of its reactants) and gradually speeds up.

Examples:

In reversible reaction,

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

reverse reaction is

$$2NH_3 \longrightarrow N_2 + 3H_2.$$

Self-assessment — 9.1

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

Solution:

Forward reaction:

As forward reaction is written from left to right, therefore,

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

Macroscopic characteristics:

In this forward reaction, N_2 and H_2 gases react to form NH_3.

Reverse reaction:

As reverse reaction is written from right to left, therefore,

$$2NH_3(g) \longrightarrow N_2(g) + H_2(g)$$

Macroscopic characteristics:

In this reverse reaction, NH_3 decomposes into N_2 and H_2.

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

Solution:

Forward reaction:

As forward reaction is written from left to right, therefore,

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

Macroscopic characteristic:

In this forward reaction H_2 and I_2 react and produce HI.

Reverse reaction:

As reverse reaction is written from right to left, therefore,

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

Macroscopic characteristic:

In this reverse reaction, HI decomposes to form H_2 and I_2

Law of Mass Action

Introduction:
Two chemists C.M Guldberg and P.Wagge in 1864 ~~experim~~ established the law of mass action to describe equilibri-um state.

Statement:
It states that "the rate at which a substance reacts is directly proportional to its active mass. The rate at which a reaction proceeds is directly proportional to the product of active masses of reactants".

Active mass:
The term active mass denotes the concentration of reactants and products at equilibrium ~~in~~ $mol.dm^{-3}$. It is represented by square brackets [ ].

Derivation of Kc expression:
Consider a hypothetical reaction in which **a** moles of reactant **A** and **b** moles of reactant **B** react to give **c** moles of product **C** and **d** moles of product **D**.

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

According to Law of mass action,

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

Rate of forward reaction $= K_f[A]^a[B]^b$

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

Rate of reverse reaction $= K_r[C]^c[D]^d$

At equilibrium state,

Rate of forward reaction = Rate of reverse reaction

$$K_f[A]^a[B]^b = K_r[C]^c[D]^d$$

Where K_f and K_r are rate constants for forward and reverse reactions respectively.

On re-arranging, $\frac{K_f}{K_r} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

$$\left[K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \right]$$

Where $K_c = \frac{K_f}{K_r}$ and is known as equilibrium constant ~~expression~~ and above expression is called equilibrium constant expression.

Equilibrium Constant K_c:

Equilibrium constant K_c is defined as the ratio of the ~~conce~~ product of concentration of products to product of concentration of reactants at equilibrium. In K_c, subscript c ~~dente~~ denotes molar concentration at equilibrium.

Dependency:

Equilibrium constant K_c is independent of initial concentration of reactants but depends upon temperature.

Example # 9.2

$$CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$$

Write equilibrium constant expression.

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

Self-assessment: 9.2

1: Solution:

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

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

Rate of forward reaction $= K_f[O_2]^3$

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

Rate of reverse reaction $= K_r[O_3]^2$

At equilibrium state,

Rate of forward reaction = Rate of reverse reaction

$$K_f[O_2]^3 = K_r[O_3]^2$$

On rearranging,

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

$$\left[K_c = \frac{[O_3]^2}{[O_2]^3} \right]$$

Where $K_c = \frac{K_f}{K_r}$ and is known as equilibrium constant and above equation is called equilibrium constant expression.

2:

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

c) $K_c = \frac{[H_2][F_2]}{[HF]^2}$

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

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

Review Ex:

5:

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

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

iii) $K_c = \frac{[Co][Cl_2]}{[CoCl_2]}$

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

Peshawar Board

Finding values of Kc:

Example 9.3 pg # 14

In the equilibrium mixture the concentration of hydrogen and iodine is 0.04 moles per dm^3 each while that of hydrogen iodide is 0.08 moles per dm^3. Find K_c of following reaction.

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

Given data:

Active mass of $H_2 = [H_2] = 0.04\ mol.dm^{-3}$

Active mass of $I_2 = [I_2] = 0.04\ mol.dm^{-3}$

Active mass of $HI = [HI]^2 = (0.08)^2\ mol.dm^{-3}$

To find:

$K_c = ?$

Solution:

K_c expression for give reaction is

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

Putting values,

$$K_c = \frac{[0.08]^2}{[0.04][0.04]}$$

$$K_c = \frac{6.4 \times 10^{-3}}{1.6 \times 10^{-3}}$$

~~$K_c = 4$ mol.dm~~

$K_c = 4$ (In this case, there are no units of K_c).

Example 9.4

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

For the above reaction, at a certain high temp. the concentrations in a particular equilibrium mixture of CO(g) is 0.0600 mol.dm^{-3}, $H_2O(g)$ is 0.120 mol.dm^{-3}, CO_2 is 0.150 mol.dm^{-3} and $H_2(g)$ is 0.3000 mol.dm^{-3}. Calculate value of K_c at this temp.

Given data:

Concentration of CO(g) = [CO] = 0.0600 mol.dm^{-3}

~~Concentration of $H_2O(g)$~~

Concentration of $H_2O(g)$ = $[H_2O]$ = 0.120 mol.dm^{-3}

Concentration of $CO_2(g)$ = $[CO_2]$ = 0.150 mol.dm^{-3}

Concentration of $H_2(g)$ = $[H_2]$ = 0.300 mol.dm^{-3}

To find: $K_c = ?$

Solution:

The K_c expression for this reaction is,

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

Putting values,

$$K_c = \frac{[0.150][0.300]}{[0.0600][0.120]}$$

$$\boxed{K_c = 6.25}$$ (There are no units of K_c in this case)

Practice problem 9.4

An equilibrium mixture of N_2, O_2 and NO gases at 1500K is determined to consist of $6.4 \times 10^{-3}\ mol.dm^{-3}$ of N_2, $1.7 \times 10^{-3}\ mol.dm^{-3}$ of O_2 and $1.1 \times 10^{-5}\ mol.dm^{-3}$ of NO. What is equilibrium constant for the system at this temp?

$$2NO_{(g)} \underset{}{\overset{1500K}{\rightleftharpoons}} N_{2(g)} + O_{2(g)}$$

Given data:

Concentration of $N_2 = [N_2] = 6.4 \times 10^{-3}\ mol.dm^{-3}$

Concentration of $O_2 = [O_2] = 1.7 \times 10^{-3}\ mol.dm^{-3}$

Concentration of $NO = [NO] = 1.1 \times 10^{-5}\ mol.dm^{-3}$

To find:

Equilibrium constant $= K_c = ?$

Solution:

The K_c expression for this reaction is,

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

Putting values,

$$K_c = \frac{[6.4 \times 10^{-3}][1.7 \times 10^{-3}]}{[1.1 \times 10^{-5}]^2}$$

$$K_c = \frac{10.88 \times 10^{-6}}{1.21 \times 10^{-10}}$$

$$K_c = 8.99 \times 10^{-6+10}$$

$$\boxed{K_c = 8.99 \times 10^{4}}$$

Exercise (Peshawar Board)

6: At equilibrium, a mixture of N_2, H_2 and NH_3 gas at 500°C is determined to consist of 0.602 mol.dm^{-3} of N_2, 0.420 mol.dm^{-3} of H_2 and 0.113 mol.dm^{-3} of NH_3. What is the equilibrium constant for this reaction.

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

Given data:

Concentration of $N_2 = [N_2] = 0.602$ mol.dm^{-3}

Concentration of $H_2 = [H_2] = 0.420$ mol.dm^{-3}

Concentration of $NH_3 = [NH_3] = 0.113$ mol.dm^{-3}

To find:

Equilibrium constant $= K_c = ?$

~~Formula~~:

Solution:

The K_c expression for this reaction is given by,

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

Putting values,

$$K_c = \frac{[0.113]^2}{[0.602][0.420]^3}$$

$$K_c = \frac{0.012}{(0.620)(0.074)}$$

$$K_c = 0.281\ dm^6/mol^2$$

Q: A reaction between gaseous SO_2 and $O_2(g)$ to produce $SO_3(g)$ takes place at 600°C. At this temp, the conc. of SO_2 is found to be 1.50 mol dm^{-3}, concentration of O_2 is 1.25 mol dm^{-3} and concentration of of SO_3 is 3.50 mol.dm^{-3}. Using the balance chemical equation, calculate K_c for this system.

Given data:

Concentration of $SO_2(g) = [SO_2] = 1.50$ mol.dm^{-3}

Concentration of $O_2(g) = [O_2] = 1.25$ mol.dm^{-3}

Concentration of $SO_3(g) = [SO_3] = 3.50$ mol.dm^{-3}

To find:

Equilibrium constant $K_c = ?$

Solution:

Balanced chemical equation for this reaction is

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

The K_c expression for this reaction is

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

Putting values,

$$K_c = \frac{[3.50]^2}{[1.50]^2[1.25]}$$

$$\boxed{K_c = 4.36\ dm^3.mol^{-1}}$$

Importance of Equilibrium constant K_c:

Equilibrium constant K_c can be used to:

1) predict the direction of chemical reaction.

Explanation:

Direction of a chemical reaction can be predicted by equilibrium constant K_c expression i.e.,

$$K_c = \frac{[\text{products}]}{[\text{reactants}]}$$

Direction of a reaction at any particular time can be predicted by comparing ratio [products]/[reactant] to value of K_c. It leads to three possibilities.

1) If ratio [products]/[reactants] is less than K_c then reaction isn't at equilibrium and more ~~reactants~~ products are required to reach equilibrium. Thus, reaction will move in forward direction until equilibrium is reached.

2) If ratio [products]/[reactants] is greater than K_c, then reaction isn't at equilibrium and more ~~products are req~~ reactants are required to reach equilibrium. Thus, reaction will move in reverse direction until equilibrium is achieved.

3) If the ratio [products]/[reactants] is equal to K_c, then reaction is at equilibrium.

ii) predict extent of chemical reaction.

Explanation:

The extent of chemical reaction depends upon magnitude of K_c:

a) If $K_c < 1$, reaction lies towards reactants which means there are more reactants and less products in equilibrium mixture.

b) If $K_c > 1$, reaction lies towards products which means there are more products and less reactants in equilibrium mixture.

c) If $K_c = 1$, ~~both reactants and products~~ equilibrium lies midway between reactants and products.

Think-Tank:

1:

i) The balanced chemical equation for this reaction is,

$$2BrCl \rightleftharpoons Br_2 + Cl_2$$

ii) K_c expression for this reaction is,

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

iii) From K_c expression,

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

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

For this reaction, K_c has no units because, total number of moles of products is equal to total number of moles of reactant.

2:

ans: Chemical equation for this reaction is

$$N_2O_4 \rightleftharpoons 2NO_2$$

K_c expression for this reaction is,

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

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

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

In this case, K_c has units because total number of moles of reactant are not equal to total number of moles of product

4)

ans: Cobalt chloride ($CoCl_2$) can definitely be used as a test for water because both anhydrous cobalt chloride and hydrated cobalt chloride have different colors. If anhydrous cobalt chloride changes its blue color and turns into pink crystals, it indicates presence of water.

3:

i) $C(s) + O_2(g) \longrightarrow CO_2(g)$

This is an irreversible reaction. This means, in this reaction, products can't form original reactants. Thus, when reaction appears to be complete, only products will be found.

ii) $2HF(g) \rightleftharpoons H_2(g) + F_2(g)$

This is a reversible reaction which means that in this reaction products convert into reactants and reactants form products simultaneously. Thus when reactions appear to be complete, there will still be some reactant alongwith products.

PYQs

Long Qs: Write down macroscopic characteristics of a reversible chemical reaction.

Reversible reaction:

The reaction in which products can react together to reform original reactants is called reversible reaction.

Macroscopic characteristics:

A reversible reaction proceeds in forward as well as reverse direction. It never goes completion. It can reach equilibrium state under suitable conditions.

Macroscopic characteristics of forward reaction:

1: In forward reaction, reactants produce products.

2: ~~Ex ex~~ Its rate is fastest in beginning and

then slows down.
3: It is written from ~~right to le~~. left to right.

Macroscopic characteristics of reverse reaction:

1: In reverse reaction, products produce reactants.
2: Its rate is zero in beginning and then speeds up.
3: It is written from right to left.

Model Paper Qs

How has Le-Chatlier's principle made it possible to get maximum amount of product from Haber's process? Write its three conditions.

ans: Le-Chatlier's principle has enabled us to obtain maximum amount of ~~pr~~ products from reversible reactions. In the case of Haber's process, N_2 and H_2 gases react to produce NH_3.

~~$N_2(g) + 3H_2(g) \rightleftharpoons$~~

$$N_2(g) + 3H_2(g) \xrightleftharpoons[Fe]{200\,atm,\ 400^{\circ}C} 2NH_3$$

Under normal conditions, only 33% of NH_3 is produce at equilibrium. But by applying this principle, we can increase this yield to 98%. Temp. is decreased while pressure is increased which increases, yield of NH_3.