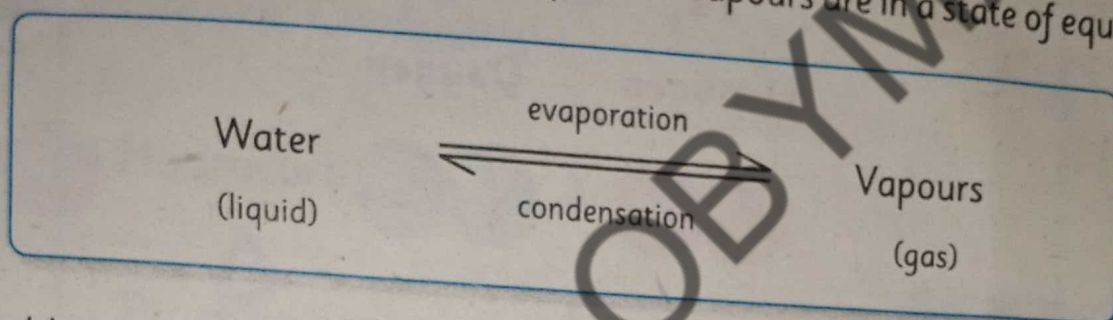


Unit	Teaching (Periods)	Assessment (Periods)	Weightage %
9 Chemical Equilibrium	8	3	10

Introduction

When liquid water is placed in a closed container at constant temperature, part of the liquid evaporates. As water begins to evaporate, at the same time, some of the vapours also begin to condense. Although the rate of evaporation is more than the rate of condensation in the start, as time passes, it becomes equal to the rate of condensation. When this happens, we may say that the liquid and vapours are in a state of equilibrium.



The double arrow shows that at the equilibrium state, both the forward and reverse processes are occurring at the same rate. Due to equilibrium, the concentration of water and vapours becomes constant and does not change with time, as long as the temperature remains constant. If the conditions of the reaction are changed, a change in the system may be observed.

Many reactions proceed to a certain extent and then apparently stop. The reactants are not completely consumed. Instead we obtain an equilibrium mixture containing both the reactants and products. When a reaction reaches this stage, it is said to have attained equilibrium.

In this unit, you will learn how and why a chemical system attains equilibrium and the characteristics of equilibrium. What happens to a system at equilibrium when a change is imposed on the system? In particular, we will discuss how to calculate the concentrations of the reactants and products present for a given system at equilibrium.

Real world reading link

Imagine a tug-of-war between two teams. As the rope between them is not moving, it might seem that neither team is pulling. In fact, both teams are pulling, but the force exerted by them is equal and opposite, so they are in complete balance or equilibrium condition.

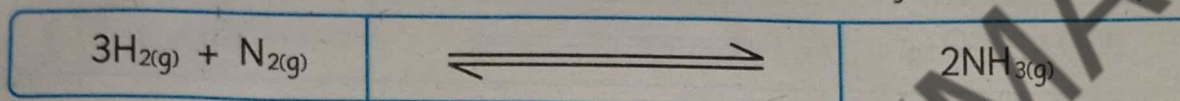
Activity 9.1

1. Perform the Tug-of-war rope activity in students.
2. On this basis, convey the concept of dynamic equilibrium to students.



9.1 Reversible Reaction and Dynamic Equilibrium

Whenever a chemical change occurs, chemical reaction takes place. In a chemical reaction, the substances that combine are called reactants and the substances formed are called products. For example, hydrogen and nitrogen (reactants) reacts to form ammonia (product).



There are a number of reactions in which the products do not recombine to form the reactants, such reactions are known as irreversible reactions. These reactions once completed cannot be reversed easily. In these reactions, the reactants and products are separated by putting a single arrow (————→) which is directed from reactants towards the products. For example, when magnesium reacts with excess of hydrochloric acid, the reaction stops when all the magnesium is used up. The products cannot be converted back to the reactants. This reaction is an example of irreversible reaction.



Beside this, there are also a number of reactions in which the products can react again to form the reactants, such reactions are known as reversible reactions. These reactions never achieve completion. These reactions can be easily reversed. In these reactions, the reactants and products are separated by putting a double half-headed arrow ($\rightleftharpoons$) between the reactants and products. This show that the reaction can move in both forward and reverse direction depending upon conditions.

For example, when mercury oxide is heated in a closed test tube from which neither the mercury nor the oxygen can escape, it decomposes and produces mercury and oxygen. Mercury and oxygen can recombine to form mercury oxide again. Thus, both forward and reverse reactions can proceed at the same time and equilibrium is achieved. Thus, at equilibrium, mercury and oxygen will react to form mercury oxide at the same rate at which mercury

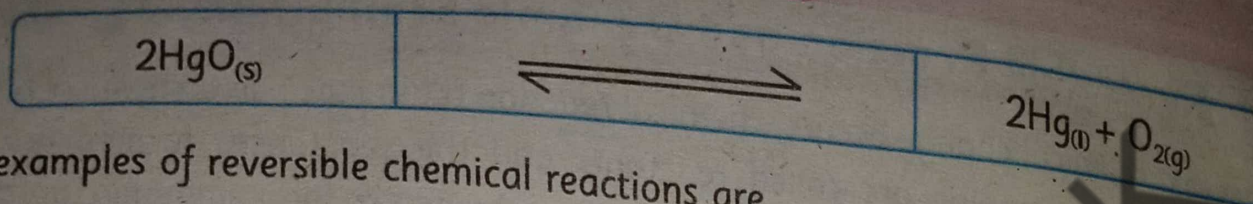


Fig.9.1 Reversible chemical process of Mercury oxide

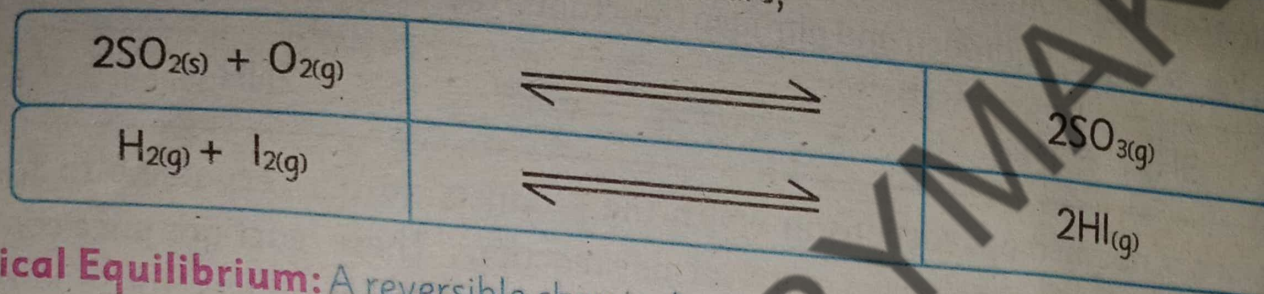
oxide decomposes into mercury and oxygen. At this point, a state of dynamic equilibrium has been reached between the two chemical reactions. Both reactions continue but there is no net change in the composition of the system.

Reading Check

Explain the meaning of double half-headed arrow in chemical equation.



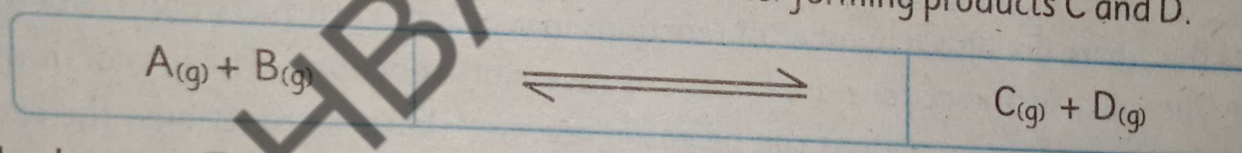
Other examples of reversible chemical reactions are,



Chemical Equilibrium: A reversible chemical reaction is in chemical equilibrium state when the rate of forward reaction equals the rate of reverse reaction and the concentrations of its products and reactants remain constant.

9.1.1 Graphical Presentation of Dynamic Equilibrium and its Examples

In a reversible reaction, dynamic equilibrium is achieved before the completion of the reaction. It is represented graphically in fig.9.2. Let us consider a general reaction in which A reacts with B in the gaseous state in a closed container forming products C and D.



Initially the concentration of reactants A and B is taken the same and is maximum and the concentration of products C and D is zero. As the forward reaction proceeds, the concentration of reactants (A and B) decreases and the concentration of products (C and D) increases simultaneously. With the passage of time, forward reaction slows down and reverse reaction speeds up. After some time, the rate of forward reaction and reverse reaction becomes equal; at this state the concentration of reactants and products becomes constant and attain equilibrium. Eventually, both reactions (forward and reverse) attain the same rate; it is called dynamic equilibrium state.

Rate of forward reaction = Rate of reverse reaction

Thus, dynamic equilibrium is an equilibrium in which the two chemical processes continue at an equal rate in opposite direction. The ratio between the concentration of products and reactants remains constant.

The examples of dynamic equilibrium are

$\text{N}_{2(g)} + 3\text{H}_{2(g)}$	$\rightleftharpoons$	$2\text{NH}_{3(g)}$
$2\text{HgO}_{(s)}$	$\rightleftharpoons$	$2\text{Hg}_{(l)} + \text{O}_{2(g)}$

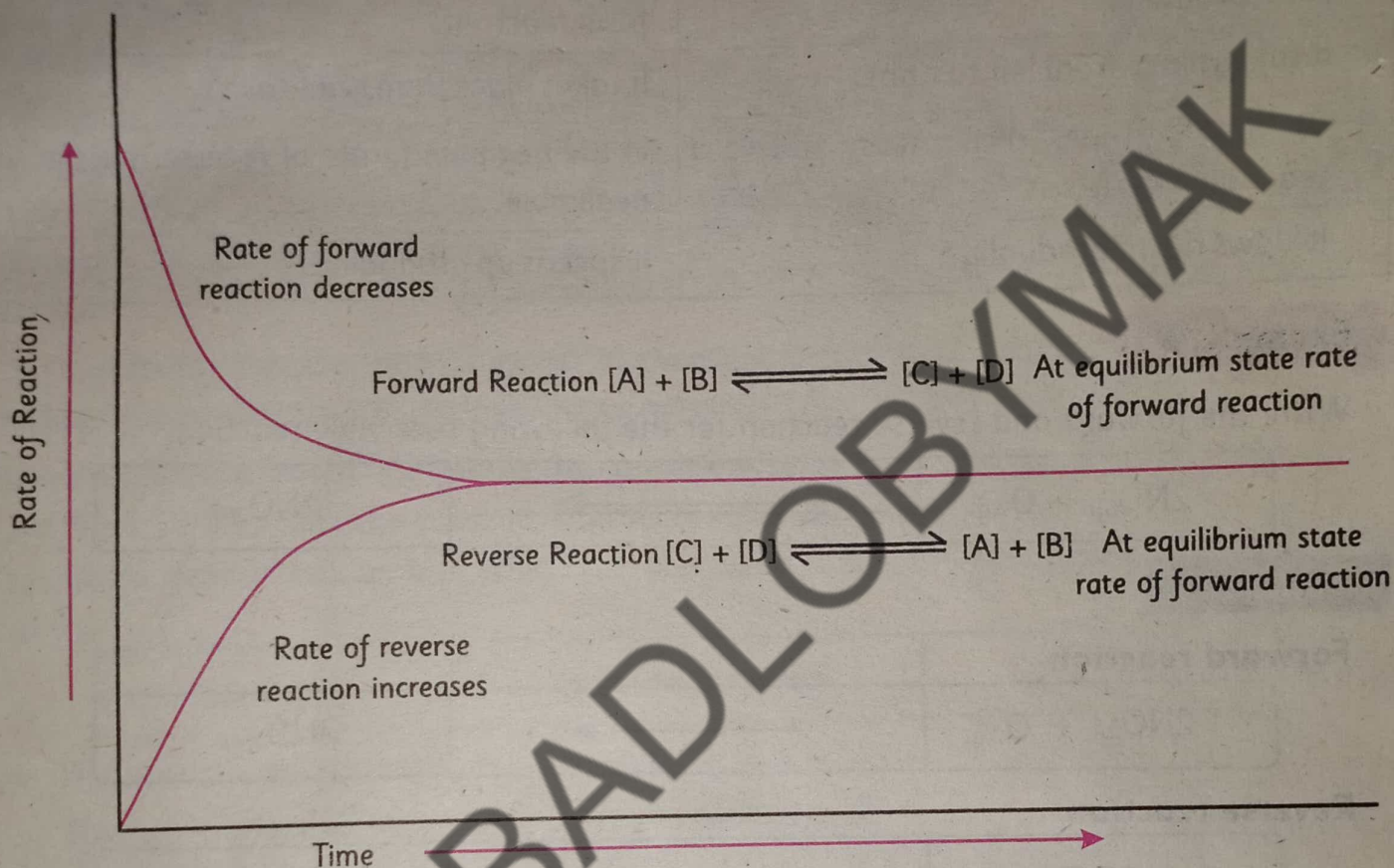


Fig.9.2 Graph showing the rate of forward and reverse reaction and equilibrium state

As long as we do not change the conditions of the system, such as pressure, temperature, concentration etc of the gases, the system will remain in equilibrium.

Reading Check

Discuss the slopes of the graphs.

Forward Reaction

It is defined as the conversion of reactants into products per unit time or the rate of chemical reaction taking place in forward direction.

Reverse Reaction

It is defined as the conversion of products back into reactants per unit time or the rate of chemical reaction taking place in reverse direction.

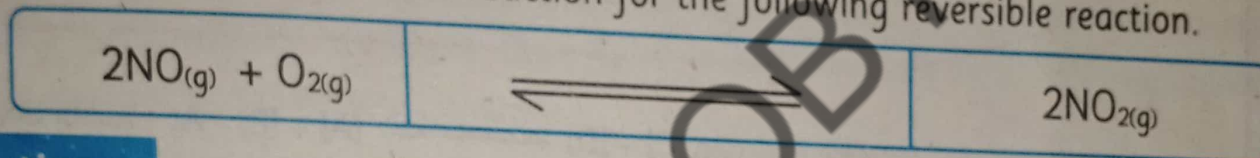
The comparison of macroscopic characteristics of forward and reverse reactions is given in table 9.1.

Table 9.1 Macroscopic Characteristics

Forward Reaction	Reverse Reaction
i. It is a reaction in which reactants react to form products.	It is a reaction in which products react to form reactants.
ii. It takes place from left to right.	It takes place from right to left.
iii. At initial stage, the rate of forward reaction is very fast.	In the beginning, rate of reverse reaction is negligible.
iv. It slows down gradually.	It speeds up gradually.

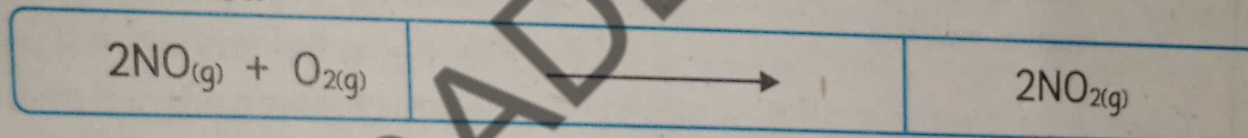
Example 9.1

Write the forward and reverse reaction for the following reversible reaction.

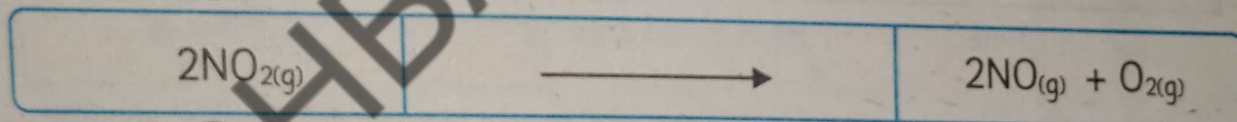


Solution

Forward reaction

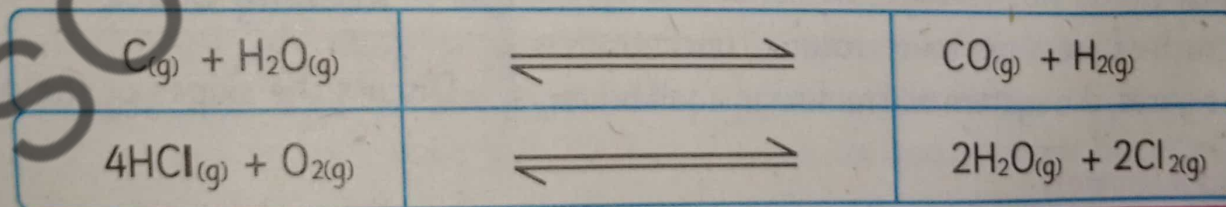


Reverse reaction



Practice problem 9.1

Write the forward and reverse reaction for the following reversible reaction.



Self Assessment

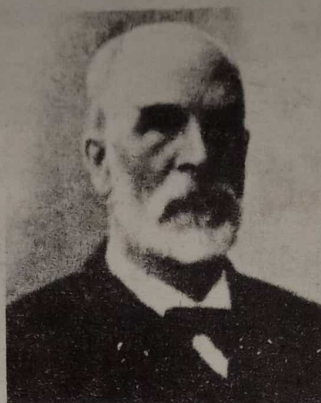
1. What is chemical equilibrium? Explain in your own words with a suitable example
2. What is dynamic equilibrium? Explain by using examples from real life.
3. Write the equilibrium equation between hydrogen and chlorine molecules as reactants and hydrochloric acid as the product.

9.2

Law of Mass Action and Derivation of the Expression for the Equilibrium Constant

9.2.1 The Law of Mass Action

In 1864, two Norwegian chemists, Guldberg and Peter Waage proposed the **Law of Mass Action** as a general description of the equilibrium condition after observing many chemical reactions. It gives a relation between the concentrations of reactants and products at equilibrium in a chemical reaction.



Cato Maximilian Guldberg
(1836-1902)



Peter Waage
(1833-1900)

This law states that **the rate or speed of chemical reaction is directly proportional to the product of the molar concentrations (active masses) of the reacting substances.**

The term **“active mass”** represents the concentration of reactants and products in mol.dm^{-3} for a dilute solution and is expressed in terms of square brackets [].

Consider a general reaction, in which two reactants A and B react to form the products C and D.



Let, The active mass or molar concentration of A is = [A]

The active mass or molar concentration of B is = [B]

The active mass or molar concentration of C is = [C]

The active mass or molar concentration of D is = [D]

Now according to the law of mass action,

Rate of forward reaction is directly proportional to the product of the concentrations of the reactants A and B. Applying the law,

$$R_{(\text{Forward})} \propto [A] [B]$$

$$(R_{(\text{Forward})} = R_f)$$

$$R_f \propto [A] [B]$$

$$R_f = K_f [A] [B]$$

Where 'k_f' is the rate constant for the forward reaction.

Similarly, for rate of reverse reaction,

$$R_{(\text{Reverse})} \propto [C] [D]$$

$$(R_{(\text{Reverse})} = R_r)$$

$$R_r \propto [C] [D]$$

$$R_r = K_r [C] [D]$$

Where ' K_r ' is the rate constant for the reverse reaction.
At equilibrium state,

Rate of forward reaction = Rate of reverse reaction

Therefore, we can write,

$$\begin{aligned} R_f &= R_r \\ K_f [A] [B] &= K_r [C] [D] \quad \text{--- (1)} \end{aligned}$$

Thus, from the above considerations, it is clear that the rate of a reaction is proportional to the molar concentrations of the reactants.

Reading Check

Define the term active mass.

9.2.2 Derivation of the Expression for Equilibrium Constant of a General Reaction

The **equilibrium constant** (K_c) is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants. It can be derived by using the rate of forward reaction and rate of reverse reaction.

Rearranging the above equation (1) we get

$$\frac{K_f}{K_r} = \frac{[C] [D]}{[A] [B]} \quad \text{--- (2)}$$

Where,

$$\frac{K_f}{K_r} = K_c$$

Therefore, we can write the above equation as,

$$K_c = \frac{[C] [D]}{[A] [B]}$$

Where K_c is called the equilibrium constant. In K_c , the subscript 'c' denotes the molar concentration at equilibrium.

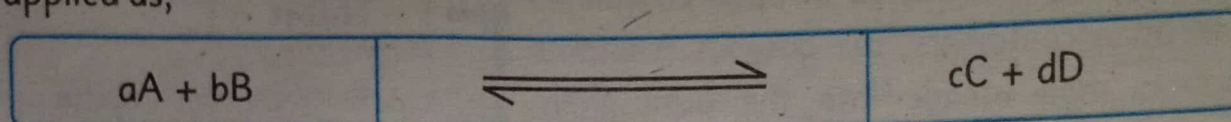
Tidbit

The value of K_c is independent of the initial concentrations. It is, however, dependent on the temperature of the system.

The general definition of the equilibrium constant can be stated as, the equilibrium constant, K_c , is the ratio of the mathematical product of the concentrations of substances formed (products) at equilibrium to the mathematical product of the concentrations of reacting substances (reactants). Each concentration is raised to the power equal to the coefficient of that substance in the chemical equation. The equation for K_c is sometimes referred to as the chemical equilibrium expression.

$$K_c = \frac{\text{Product of concentration of products raised to the power of coefficient}}{\text{Product of concentration of reactants raised to the power of coefficient}}$$

Let us consider a general reaction with different number of moles, so the law of mass action can be applied as,



Where 'a, b, c, and d' are the molar concentration of A, B, C, and D respectively.

The equilibrium constant expression can be written as,

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

A large value of K_c shows that at equilibrium the product concentrations are greater than those of the reactants and vice versa. The equilibrium constant of a reaction is independent of pressure, concentration and catalyst and its value is constant for a particular reaction.

Where ' K_c ' is called the rate constant in terms of concentrations.

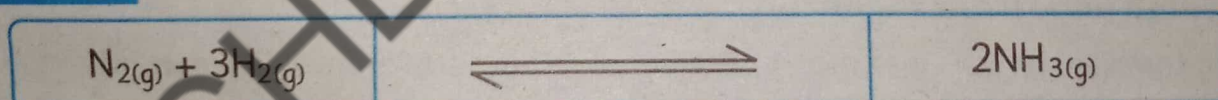
Interesting facts

The law of chemical equilibrium is sometimes known as the law of mass action. Before the term "concentration" was used, the concept of amount per unit volume was called "Active mass".

Example 9.2

Write an equilibrium constant expression for the reaction of nitrogen and hydrogen to form ammonia at 150 to 200 atmospheric pressure and 450 to 500°C while Iron is used as catalyst.

Solution



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

Practice problem 9.3

Write an equilibrium constant expression for the following reversible reactions.

a. $N_{2(g)} + 2O_{2(g)}$	$\rightleftharpoons$	$2NO_{2(g)}$
b. $2N_{2(g)} + O_{2(g)}$	$\rightleftharpoons$	$2N_2O_{(g)}$
c. $2SO_{2(g)} + O_{2(g)}$	$\rightleftharpoons$	$2SO_{3(g)}$
d. $PCl_{3(g)} + Cl_{2(g)}$	$\rightleftharpoons$	$PCl_{5(g)}$

9.2.3 Conditions for Equilibrium

Equilibrium cannot be attained in open containers because in open containers, the gaseous reactants and products may escape into the atmosphere leaving behind no possibility of attaining equilibrium. When a system reaches equilibrium, it will remain in the same state forever, if the conditions do not change. When the conditions are disturbed, the equilibrium will also be disturbed. The system always tries to regain its equilibrium. The reaction will move either in the forward direction or in the reverse direction to attain equilibrium.

As long as the conditions are kept constant, the reaction will remain in the equilibrium state. Conditions are following,

1. Concentration of reactants and products,
2. Temperature of the system,
3. Pressure of the system,
4. Volume of the system,
5. Catalyst, if used in the system remain unchanged.

Tidbit

At higher temperature, the particles have more kinetic energy, they move at a faster speed, and there will be more successful collisions resulting in the formation of more products. Whereas, the low temperature, in the refrigerator slows down the decomposition reaction.

Society, Technology And Science

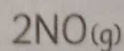
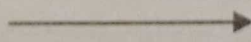
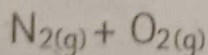
Explain how components of the atmosphere can be used successfully in producing important chemicals.

Except for helium, which is mostly extracted from natural gas. Oxygen, nitrogen and other rare gases are extracted from the air that makes up Earth's atmosphere.

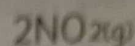
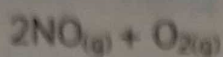
Nitrogen and oxygen are important components of atmosphere. Air is mainly composed of nitrogen, oxygen, carbon dioxide and inert gases. Nitrogen (78%) and oxygen (21%) constitute 99% of the atmosphere.

By fractional distillation, the air is separated into different fractions. These fractions are used in manufacturing of different chemicals.

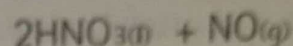
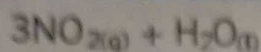
Nitrogen: In Birkland – Eyde process, nitrogen produces nitric acid, which is an important chemical compound. During this process, nitrogen reacts with oxygen and forms nitric oxide, NO.



The nitric oxide is further oxidized into NO_2 in the presence of excess of oxygen.



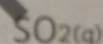
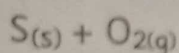
The NO_2 gas is dissolved in water and nitric acid is formed.



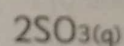
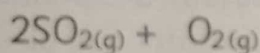
In Haber process, nitrogen reacts with hydrogen and produces ammonia (NH_3). Ammonia acts as a raw material in the manufacturing of fertilizers.

Oxygen: It also plays an important role and produces numerous chemicals, such as oxides of nitrogen, which ultimately produce nitric acid. Nitric acid is used in the production of fertilizers.

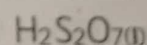
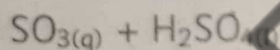
It also reacts with sulphur to form sulphur dioxide (SO_2).



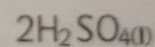
The sulphur dioxide (SO_2) is further oxidized into sulphur trioxide (SO_3) in excess of oxygen.



Sulphur trioxide (SO_3) is dissolved in sulphuric acid, which produce oleum (pyrosulphuric acid).



Oleum is dissolved in calculated amount of water and concentrated sulphuric acid is obtained.



Sulphuric acid plays basic role in most chemical reactions, and is considered the king of chemicals.

9.2.4 Ways to Recognize Equilibrium

In order to recognize the equilibrium state of chemical reaction, the following methods can be used:

- Physical methods** such as Refractometry, Polarimetry, Spectrophotometry etc.
- Chemical methods** such as titration etc.

In both methods, the equilibrium state of chemical reaction can be recognized by knowing the concentration of reactants and products at regular intervals of time. When the concentration of reactants and products are observed constant, then the reaction is said to be in equilibrium state.

Tidbit

A catalyst is a substance that changes the rate of a chemical reaction but remains chemically unchanged itself.

Self Assessment

1. Define the law of mass action.
2. What is equilibrium constant for a chemical reaction? How is it written?
3. Write down necessary conditions for equilibrium.

9.3 Equilibrium Constant and Its Units

Equilibrium constant is the ratio of the product of the concentration of the products to the product of the concentration of the reactants at equilibrium state.

The numerical value of K_c for a particular equilibrium system is obtained experimentally. The chemists examine the equilibrium mixture and determine the concentration of all substances. When we know the balanced chemical equation of a reversible reaction we can write the equilibrium constant (K_c). While writing the equilibrium constant, the products side is written as numerator and the reactants as denominator. Thus we can calculate the numerical value and unit of K_c , by putting the equilibrium values of the products side at numerator and the reactant side at denominator and the coefficient is raised to the power of that substance in the chemical equation.

Equilibrium Constant

The square brackets show concentration in mol.dm^{-3}

"C" shows that it is in terms of concentration

The exponents are the numbers in front of each substance in the chemical equation

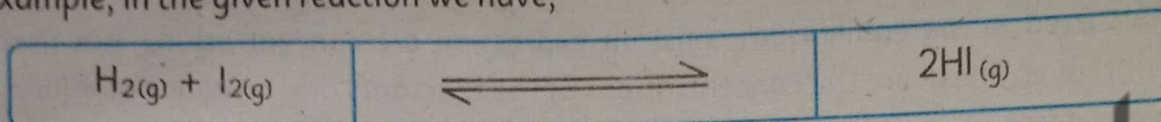
Right-hand (products) side of the equation on top; left-hand (Reactants) on the bottom

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

Fig.9.3 Equilibrium constant, K_c

The unit of K_c depends on the equilibrium constant expression for the given reaction.

- i. K_c has no units for a reaction in which the number of moles of reactants and products are equal. This is because concentration units cancel out in the expression for K_c . For example, in the given reaction we have,

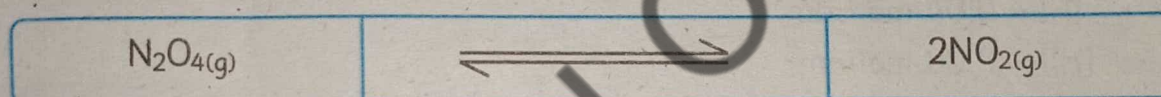


$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$\text{Units of } K_c = \frac{[\text{mole.dm}^{-3}]^2}{[\text{mole.dm}^{-3}][\text{mole.dm}^{-3}]} = \text{No Unit}$$

- ii. K_c has units for a reaction in which the number of moles of products is greater than reactants in a balanced chemical equation.

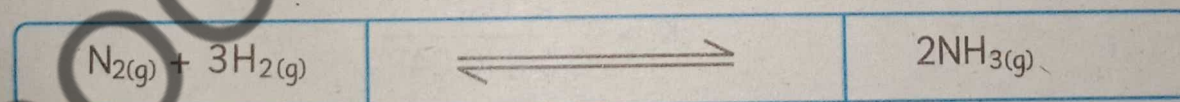
For example, in the given reaction we have,



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

$$\text{Units of } K_c = \frac{[\text{mole.dm}^{-3}]^2}{[\text{mole.dm}^{-3}]} = \text{mole.dm}^{-3}$$

- iii. Similarly, K_c has units for a reaction in which the number of moles of products is less than the reactants in a balanced chemical equation. For example, in the given reaction we have,



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

$$\text{Units of } K_c = \frac{[\text{mole.dm}^{-3}]^2}{[\text{mole.dm}^{-3}][\text{mole.dm}^{-3}]^3} = \frac{1}{[\text{mole.dm}^{-3}]^2}$$

It is clear from the above equations that the units of K_c depend upon the equation. It varies from equation to equation.

Finding Values of Equilibrium Constants

The value of equilibrium constant can be calculated if we know the concentrations of the reactants and products in the equilibrium mixture.

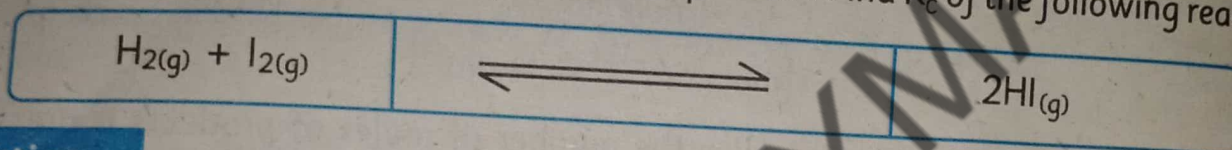
The values used in the equilibrium constant expression are concentrations, not numbers of moles. Moreover, these are the concentrations of the reactants and products at equilibrium.

Reading Check

In which case the K_c has no value?

Example 9.3

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 the following reaction.



Solution

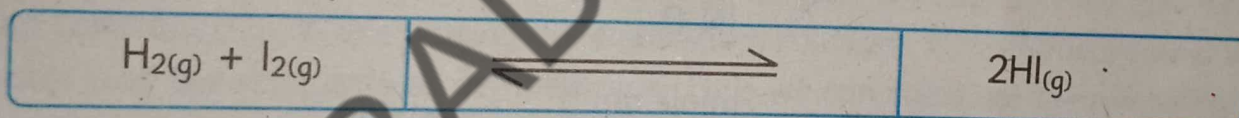
Given: $[\text{H}_2] = 0.04 \text{ mol/dm}^3$

$[\text{I}_2] = 0.04 \text{ mol/dm}^3$

$[\text{HI}] = 0.08 \text{ mol/dm}^3$

$K_c = ?$

The balanced chemical equation is



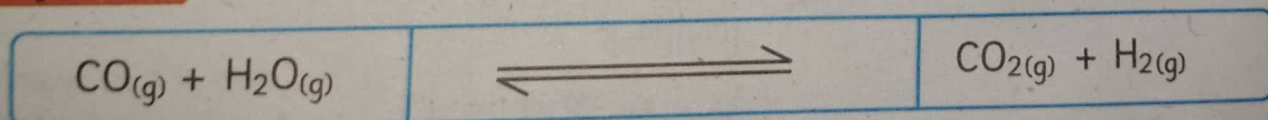
The chemical equilibrium expression is $K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$

Substitute the given values for the concentrations into the equilibrium expression.

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

$$K_c = 4$$

Example 9.4



For the above reaction at a certain high temperature, the concentrations in a particular equilibrium mixture of $\text{CO}_{(g)}$ is 0.0600 mol/dm^3 , $\text{H}_2\text{O}_{(g)}$ is 0.120 mol/dm^3 , CO_2 is 0.150 mol/dm^3 and $\text{H}_{2(g)}$ is 0.300 mol/dm^3 . Calculate the value of the equilibrium constant at this temperature.

Solution

Given: $[\text{CO}] = 0.0600 \text{ mol/dm}^3$
 $[\text{H}_2\text{O}] = 0.120 \text{ mol/dm}^3$
 $[\text{CO}_2] = 0.150 \text{ mol/dm}^3$
 $[\text{H}_2] = 0.300 \text{ mol/dm}^3$
 $K_c = ?$

The equilibrium constant expression for this reaction is

$$K_c = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

Substituting the equilibrium concentrations of reactants and products the value of the equilibrium constant:

$$K_c = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$K_c = \frac{[0.150 \text{ mol/dm}^3][0.300 \text{ mol/dm}^3]}{[0.0600 \text{ mol/dm}^3][0.120 \text{ mol/dm}^3]} = 6.25$$

Practice problem 9.4

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

9.4 Importance (applications) of Equilibrium Constant

The value of equilibrium constant is specific and remains constant at a particular temperature. The value of K_c helps us to predict.

1. Direction of Reaction

We can determine the direction of reaction with the help of equilibrium constant expression.

$$K_c = \frac{[\text{Products}]}{[\text{Reactants}]}$$

The direction of chemical reaction at any particular time can be predicted by means of $[\text{products}] / [\text{reactants}]$ ratio. The value of $[\text{products}] / [\text{reactants}]$ ratio leads to one of the following three possibilities.

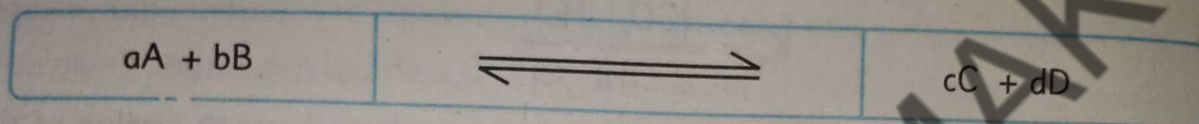
- a. When the ratio of $[\text{products}] / [\text{reactants}]$ is less than K_c . The system is not at equilibrium and more products are required to reach the equilibrium. Therefore, the reaction will move in the forward direction until equilibrium is reached.

- b. When the ratio of [products] / [reactants] is greater than K_c . The system is not at equilibrium and more reactants are required to reach equilibrium. The reaction will move in the reverse direction until equilibrium is reached.
- c. When the ratio of [products] / [reactants] is equal to K_c . The system is at equilibrium.

2. The Extent of a Chemical Reaction

Extent to which a reaction may proceed can be determined with the help of equilibrium constant expression.

Consider the general reaction,



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

The extent of reaction depends upon the magnitude of ' K_c ' so when,

i. K_c is Very Small

When the concentration of [A] and [B] is large and that of [C] and [D] is small, the equilibrium mixture will contain the reactants mainly and only a small amount of products will be present. It reflects that the reaction does not proceed appreciably in the forward direction.

ii. K_c is very large

When the concentration of [A] and [B] is small and that of [C] and [D] is large, the equilibrium mixture will consist almost entirely of products and a small amount of reactants will be present. This indicates that the reaction is completed in forward direction.

iii. K_c is Neither Very Large Nor Very Small

The concentration of the products and reactants will be very close and hence, K_c is close to 1.0. Thus, equilibrium mixture will contain appreciable amount of products and reactants.

3. The Effect of External Conditions on the Position of Equilibrium

When a system reaches equilibrium it will remain in the same state indefinitely, if the conditions do not change. However, the equilibrium state of a system is disturbed if external conditions are changed, e.g. change of pressure, temperature and concentrations of reactants and products alter the position of the equilibrium. Whenever, the equilibrium is disturbed by changes in the external conditions, the system always tends to restore equilibrium.

Point to Ponder!

Does a small value of K_c indicate a slow reaction? Use an example to justify your answer.

Key Points

- The reactions in which the products do not recombine to form the reactants are called irreversible reactions.
- The reactions in which the products can react to form the reactants again are called reversible reactions. These reactions do not reach to completion.
- In equilibrium state, the rate of its forward reaction equals the rate of its reverse reaction and the concentrations of its products and reactants remain constant.
- A reaction is at dynamic equilibrium if the rate of the forward reaction is equal to the rate of the reverse reaction.
- The conversion of reactants into products per unit time or the rate of chemical reaction taking place in forward direction is called forward reaction.
- The conversion of products back into reactants per unit time or the rate of chemical reaction taking place in reverse direction is called reverse reaction.
- According to the law of mass action, the rate of chemical reaction is directly proportional to the product of the molar concentrations (active masses) of the reacting substances.
- Active mass means the concentration of reactants and products in mol.dm^{-3} for a dilute solution.
- Equilibrium constant K_c is the product of the equilibrium concentrations of the products divided by the product of the equilibrium concentrations of the reactants, with each concentration term raised to a power equal to the coefficients of the substance in the balanced equation.
- A very large value of K_c indicates that the reaction is almost complete. If the value of K_c is very small, then the reaction proceeds slowly in the forward direction.
- There are two ways to recognize equilibrium i.e. physical and chemical. The equilibrium constant can be used to predict the direction and extent of the chemical reaction.

Exercise

A. Choose the Correct Option.

- At dynamic equilibrium
 - The reverse reaction stops
 - The forward reaction stops
 - Both forward and reverse reactions stop
 - Both forward and reverse reactions continue at the same rate
- A reversible reaction have the following characteristics except
 - They proceeds in both direction
 - They never complete
 - Products do not form reactants again
 - They are represented by $\rightleftharpoons$
- The equilibrium constant (K_c) expression for the given reaction is,

$N_{2(g)} + O_{2(g)}$	$\rightleftharpoons$	$2NO_{(g)}$
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 - $\frac{[2N][2O]}{[2NO]}$
 - $\frac{[N]^2 [O]^2}{[2NO]}$
 - $\frac{[N_2] [O_2]}{[NO]^2}$
 - $\frac{[NO]^2}{[N_2] [O_2]}$
- The equilibrium constant (K_c) is:
 - The sum of the two reactants
 - The difference of the two rate constants
 - The ratio of the two rate constants
 - The product of the two rate constants
- When the value of K_c is very small it shows that
 - reaction will go in the forward direction
 - reaction will go in the reverse direction
 - reaction is at equilibrium
 - equilibrium will never establish
- For which reaction the K_c will have no units?

a.	$4NH_{3(g)} + 5O_{2(g)}$	$\rightleftharpoons$	$4NO_{(g)} + 6H_2O$
b.	$N_{2(g)} + 3H_{2(g)}$	$\rightleftharpoons$	$2NH_{3(g)}$
c.	$N_{2(g)} + O_{2(g)}$	$\rightleftharpoons$	$2NO_{(g)}$
d.	$CO_{(g)} + 3H_{2(g)}$	$\rightleftharpoons$	$CH_{4(g)} + H_2O_{(g)}$

7. When the value of K_c is very large it shows that
- Reaction is at equilibrium
 - Equilibrium will never be achieved.
 - Reaction will move in the forward direction
 - Reaction will move in the reverse direction
8. Active mass means
- The total mass of reactants
 - The total mass of products
 - The total mass of products and reactants
 - Concentration of reactants and products in moles per dm^3 in a dilute solution

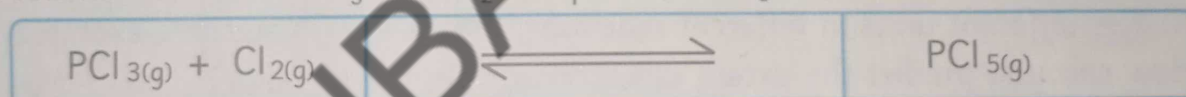
9. For a reversible reaction

$$K_c = \frac{[C]^4}{[A]^3[B]^2}$$

The equation will be,

a.	$4C_{(g)}$	$\rightleftharpoons$	$3A_{(g)} + 2B_{(g)}$
b.	$4C_{(g)}$	$\rightleftharpoons$	$A_{3(g)} + B_{2(g)}$
c.	$A_{3(g)} + B_{2(g)}$	$\rightleftharpoons$	$C_{4(g)}$
d.	$3A_{(g)} + 2B_{(g)}$	$\rightleftharpoons$	$4C_{(g)}$

10. The reaction between PCl_3 and Cl_2 will produce PCl_5 .

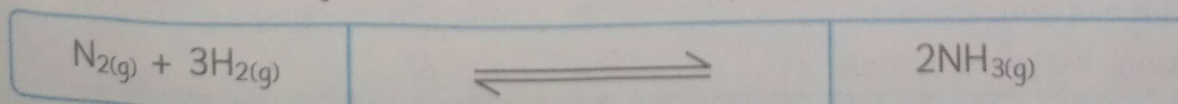


The units of K_c for this reaction are,

- a. $\text{mol} \cdot \text{dm}^{-3}$ b. $\text{mol}^{-1} \cdot \text{dm}^{-3}$ c. $\text{mol}^{-1} \cdot \text{dm}^3$ d. $\text{mole} \cdot \text{dm}^3$

B. Short questions

- Define chemical equilibrium with two examples.
- How would you identify that dynamic equilibrium is established?
- Compare the different macroscopic characteristics of forward and reverse reactions.
- What information is required to predict the direction of a chemical reaction?
- Relate the active mass with rate of a chemical reaction.
- 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 the reaction at this temperature?



Answer
0.286

7. State conditions necessary for chemical equilibrium.

8. Write equilibrium constant expression for the following reactions

$\text{N}_{2(g)} + 3\text{H}_{2(g)}$	$\rightleftharpoons$	$2\text{NH}_{3(g)}$
$2\text{H}_{2(g)} + \text{O}_{2(g)}$	$\rightleftharpoons$	$2\text{H}_2\text{O}_{(g)}$
$4\text{NH}_{3(g)} + 5\text{O}_{2(g)}$	$\rightleftharpoons$	$4\text{NO}_{(g)} + 6\text{H}_2\text{O}_{(g)}$

9. A reaction between gaseous sulphur dioxide and oxygen gas to produce gaseous sulphur trioxide takes place at 600°C . At this temperature, the concentration of SO_2 is found to be 1.50 mol/dm^3 , the concentration of O_2 is 1.25 mol/dm^3 , and the concentration of SO_3 is 3.50 mol/dm^3 . Using the balanced chemical equation, calculate the equilibrium constant for this system.

Answer 4.36

10. Describe the effect of temperature on equilibrium constant.

C. Long questions.

- $\text{SO}_{3(g)}$ decomposes to form $\text{SO}_{2(g)}$ and $\text{O}_{2(g)}$. For this reaction write,
 - Chemical equation
 - K_c expression, and
 - Derive the units of K_c for this reaction.
- (a). Describe the equilibrium state with the help of a graph and an example.
(b). Define the law of mass action.
- (a). Derive an expression for the equilibrium constant and explain its units.
(b). How can you predict the direction of reaction for the K_c value?
- (a). K_c has different units in different reactions. Prove it with suitable examples.
(b). How can you predict the extent of reaction from the K_c value?
- (a). K_c expression for a reaction is given below,

$$K_c = \frac{[\text{Cl}_2]^2 [\text{H}_2\text{O}]^2}{[\text{HCl}]^4 [\text{O}_2]}$$

For this reaction write,

- Reactants and products
 - Derive the units of K_c
- (b). Explain the importance of equilibrium constant, support your answer with examples and reasons.

Project

Ammonia is an important chemical and is used in the production of urea. Which of the following factors affect the production of ammonia (NH_3)?

- Change in temperature
- Change in pressure
- Change in concentration
- Catalyst