Chemistry Notes

Chapter: Chemical Equilibrium

Reversible and Irreversible Reactions:

We are aware that when certain raw materials, referred to as reactants, are chemically reacted, products are created. The term "irreversible reaction" refers to a reaction in which the reactants are entirely converted into products. Meaning that this reaction will favor moving forward in a single direction from the reactants' left side to the right (products side). We display irreversible reactions using single-headed arrows, like as

$$A + B \longrightarrow C + D$$

While reversible means that the products created by the reaction of reactants will react with each other and result in the synthesis of the original reactants once more. means that the reaction will move from the right (product side) to the left in the opposite direction (reactants side). Therefore, it implies that reversible reactions never come to an end and instead continue indefinitely. The reversible reaction occurs in both the forward and reverse directions at the same time. We display reversible reactions using double-headed arrows, like as

$$A + B \rightleftharpoons C + D$$

Dynamic equilibrium and reversible reactions:

Reversible reactions aid in achieving dynamic equilibrium.

What exactly is dynamic equilibrium then?

According to one definition of dynamic equilibrium, it occurs when a reaction moves simultaneously in both the forward and backward directions. Because of this, we already said that reversible processes aid in the development of dynamic equilibrium.

Dynamic equilibrium illustration:

The liquid turns into vapours when it is heated in a sealed container. The vapours will rise and will also collide with liquid surfaces, which will cause some of the vapours to condense and return to a liquid state. Thus, we can describe this system as having simultaneous evaporation (ahead reaction) and condensation (reverse reaction).

Examples of Reversible Reaction:

Example 1:

$$2SO_{2(g)} + O_2 \underset{450^\circ C,\ 200\ atm}{\overset{V_2O_5}{\rightleftharpoons}} 2SO_{3(g)}$$

The double-headed arrow between the reactants and products indicates that the aforementioned reaction is reversible. In this reaction, sulphur dioxide (the reactant) and oxygen (the reactant) combine at 450 ^{o}C and 200 atm in the presence of a vanadium pentoxide catalyst to produce sulphur trioxide. As the reactants move forward and produce products, this reaction is known as a forward reaction. Since no reaction occurs and no product is produced when the forward reactions first begin, the concentration of sulphur trioxide will be zero.

As time goes on, the forward reaction's rate will increase, and at this point, the rate of the backward reaction will be zero. Therefore, the concentration of the reactants (sulfur dioxide and oxygen) will start falling on the left side and the rate of the forward reaction will start slowing down after some time when the product (sulfur-trioxide) creation begins. As the pace of the forward reaction starts to slow down, the rate of the backward reaction will also pick up. This will result in the production of the original reactants, oxygen and sulphur dioxide, once more from the product, sulphur trioxide. Therefore, if both the forward and backward reaction rates are simultaneously equal, then the reactants and products have reached equilibrium.

Therefore, the condition of a system at which the rates of the forward and reverse reactions are equal is known as the chemical equilibrium state.

When reactant and product concentrations are equal, the chemical process is said to be in equilibrium, and both will be in balance.

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The following steps can be used to write the above reaction:

Forward Reaction:

$$2SO_{2(g)} + O_2 \xrightarrow[450^\circ C,\ 200\ atm]{V_2O_5} 2SO_{3(g)}$$

Reverse Reaction:

$$2SO_{3(g)} \xrightarrow[450^\circ C,\ 200\ atm]{V_2O_5} 2SO_{2(g)} + O_2$$

Example 2:

$$N_2 + 3H_2 \underset{400^\circ C,\ 200\ atm}{\overset{Fe}{\rightleftharpoons}} 2NH_3$$

Forward Reaction:

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

Reverse Reaction:

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

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Example 3:

$$2N_{2(g)} + O_{2(g)} \underset{\text{Spark}}{\overset{\text{Electric}}{\rightleftharpoons}} 2NO_{2(g)}$$

Forward Reaction:

$$2N_{2(g)} + O_{2(g)} \xrightarrow[\text{Spark}]{\text{Electric}} 2NO_{2(g)}$$

Reverse Reaction:

$$2NO_{2(g)} \xrightarrow[\text{Spark}]{\text{Electric}} 2N_{2(g)} + O_{2(g)}$$

Example 4:

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

Forward Reaction:

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

Reverse Reaction:

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

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Law of Mass Action:

In 1864, C.M. Gulberg and P. Waage proposed the concept of the law of mass action. To demonstrate how this law works, add some sugar to a glass of water, shake it vigorously, or stir it with a spoon. Observation will reveal that a sugar crystal's surface molecules are in direct contact with water, and the water molecules around it aid in their dissolution, causing them to break into ions that carry an electric charge.

The number of reactants that participate by dissociating into ions and carrying an electric charge may now be stated as the active mass of that sugar. The pace of any reaction depends on the active masses of the reactants involved in that reaction.

We can also define active mass as the molar concentration of the reactants and products, which we represented as mol/dm3. **Square brackets [ ]** are used to denote the active mass of material, whether it is a reactant or a product.

Consider the following response.

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

In the light of the law of mass action;

Rate of forward reaction $\propto [H_2]^1 [I_2]^1$

Rate of forward reaction = $K_f [H_2]^1 [I_2]^1$

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

Rate of reverse reaction = $Kr [HI]^2$

The number 1 and 2 in this equation represent the number of moles of reactant and product, respectively, and $[H_2]^1 [I_2]^1$ is the concentration of H_2 and I_2 in terms of mol/dm^3, while K_f and Kr are proportionality constants for forward and reverse processes, respectively.

We may set up the response as

Rate of forward reaction = Rate of reverse reaction

$$K_f [H_2]^1 [I_2]^1 = Kr [HI]^2$$

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By rearranging,

$$\frac{K_f}{Kr} = \frac{[HI]^2}{[H_2]^1\,[I_2]^1}$$

$$Kc = \frac{[HI]^2}{[H_2]^1\,[I_2]^1}$$

So, we are aware that Kc = K_f / Kr, also known as the equilibrium constant, and that the c in front of K stands for the molar concentrations of the reactants and products in the equilibrium state.

Therefore, to define the **Kc or equilibrium constant**, we can state that it is the result of the molar ratio of the products and reactants. Initial reactant concentration is not affected by Kc but the temperature is.

Equilibrium Requirements:

If you keep an eye on things once equilibrium is reached, you'll see that:

The concentration of the reactants and products will remain the same/balanced/constant

The system's volume, temperature, and pressure will all remain unchanged/balanced/constant.

Point of confusion:

The concentration of the reactants and products will be the same or constant in this case, meaning that it won't change. The same goes for the temperature, pressure, and volume. The yield of the product can then be enhanced by varying the temperature, pressure, volume, and concentration of the reactants, according to Le Chatellier's principle. You will therefore be confused by the fact that in equilibrium, which we examined, it remained unchanged or steady, however here, they are changeable and we can decrease and raise it. Therefore, always keep in mind that the beauty of an equilibrium state is that it maintains the concentration of the reactants and products at a constant level on its own. The concentration of the reactants and products must remain constant for that state to be an equilibrium state.

How to tell if a system is in equilibrium:

We are aware that the forward and reverse reaction rates will equalize at the point of equilibrium, which implies the concentrations of the reactants and products will balance out. How will we ascertain physically and chemically that the concentrations of the reactants and products have indeed become balanced or equal?

The most popular techniques for determining equilibrium states are Titration and spectroscopy.

Equilibrium constant and its units:

We are aware that Kc is the result of dividing the molar concentration of the reactants by the molar concentration of the products. The active mass in the aforementioned rule of mass action is the molar concentration of the reactants and products, which we stated as mol/dm^3. Therefore, the concentration of each reactant and product can be represented as mol/dm^3 in the square bracket [ ]. In this manner, the unit of Kc may be determined. The unit of Kc relies on the number of moles of the reactants and products, for example, so sometimes the Kc will have a unit, and other times it won't.

Here are the responses. We'll now determine their equilibrium constant (Kc) unit.

Example 1:

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

Solution:

$$Kc = \frac{[HI]^2}{[H_2]^1\,[I_2]^1}$$

$$Kc = \frac{\cancel{[mol/dm^3]^2}}{\cancel{[mol/dm^3]^1}\;\cancel{[mol/dm^3]^1}}$$

Kc = no units

Because the number of moles of reactants and products is equal in this instance, there is no unit of Kc. As a result, the two moles of products will cancel out the two moles of reactants, leaving no unit of Kc.

Example 2:

$$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$$

Solution:

$$Kc = \frac{[NO_{2(g)}]^2}{[N_2O_{4(g)}]^1}$$

$$Kc = \frac{\cancel{[mol/dm^3]^2}}{\cancel{[mol/dm^3]^1}}$$

$$Kc = mol/dm^3$$

We can see that there are two moles of product and one mole of reactants in the second scenario. As a result, to determine the unit of Kc, one mole of the product must cancel with one mole of reactant, leaving behind the second mole of reactant, which is the unit of Kc.

Example: 3

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

Solution:

$$Kc = \frac{[2NO_{2(g)}]^2}{[NO_{(g)}]^2 \quad [O_{2(g)}]^1}$$

$$Kc = \frac{\cancel{[mol/dm^3]^2}}{\cancel{[mol/dm^3]^2}\ [mol/dm^3]^1}$$

$$Kc = mol^{-1}dm^3$$

You can see that there are more moles of reactants than products in this reaction, thus the two moles of products will cancel out with the two moles of reactants, leaving the third mole of reactant behind, which will be the unit of Kc.

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Importance of equilibrium constant:

The equilibrium constant can help us understand some crucial facts about a reaction. When we don't know the equilibrium concentration of the equilibrium mixture, Kc uses the initial concentration of the reactants to determine the equilibrium concentration of the equilibrium mixture.

When a reaction is underway and we are unsure of its direction—whether it will move ahead or backward—Kc will decide the reaction's course.

When a reaction is underway and we are unsure of its magnitude or the rate at which it is progressing, Kc will determine the reaction's size or velocity.

When a reaction is in equilibrium, any changes that take place at this point in the reaction will have an impact on the reaction depending on Kc. Industrial chemists also determine the impact of changes in concentration, temperature, pressure, and other variables before determining the impact of changes that take place at equilibrium.

Le Chatellier's Principle:

According to Le Chatellier's principle, a reaction will migrate its equilibrium to either the left or the right side of the reaction when its concentration, temperature, or pressure are changed while it is in an equilibrium condition. Therefore, to synthesize more and more products, we will make the kinds of adjustments that cause the equilibrium to shift in favor of the product side.

According to this theory, **when a reaction is endothermic**, it means that it will absorb energy. As a result of the system's increased temperature as a result of the energy absorption, the reaction will then prefer the product side.

When a reaction is exothermic, however, it indicates that energy will be released; as a result, the system's temperature will drop and the equilibrium will shift to the product side.

When there are more moles on the reactant side and fewer on the product side, the equilibrium will move toward the product side when pressure is increased. This is because pressure and volume are known to be inversely related to one another. The volume reduces as pressure rises and vice versa.

The reaction will favor the product side as the concentration of the reactant's side rises.

For Example:

$$N_2 + 3H_2 \underset{400\,°C,\ 200\ atm}{\overset{Fe}{\rightleftharpoons}} 2NH_3$$

Δ H = -92.4 kJ

In this response, we can observe that there are more moles on the left side and fewer on the right. Therefore, by raising the pressure, we know that the reaction will move from the volume side to the product side, increasing the yield of ammonia.

Furthermore, we would use a low temperature to produce more product because we know that the reaction is exothermic and will release heat or energy; otherwise, if we use a high temperature, the reaction will favor moving oppositely, producing reactants once more.

Example 2:

$$2SO_2 + O_2 \xrightarrow[400\ °C,\ 200\ atmp]{V_2O_5} 2SO_3$$

Δ H = -196 KJ

We will use high pressure in this reaction because we can see that the number of moles is greater on the reactant side and lower on the product side. Since pressure and volume are inversely proportional, as pressure is increased, the equilibrium will move to the side with the smaller volume (the product side), giving us a high yield of Sulphur trioxide.

Furthermore, we would use a low temperature to produce more product because we know that the reaction is exothermic and will release heat or energy; otherwise, if we use a high temperature, the reaction will favor moving oppositely, producing reactants once more.