# EXPERIMENT

AIM

To study the effect of concentration on the equilibrium between ferric ions and thiocyanate ions.

MATERIAL REQUIRED

Test tubes, test tube stand, droppers, glass rods, beakers, weight box, measuring flask and measuring cylinders. Ferric chloride (0.1 M), potassium thiocyanate (0.1 M) and potassium chloride (0.1 M).

THEORY

When a system in equilibrium is suddenly disturbed, it will respond in some way until equilibrium is re-established. Consider the equilibrium between ferric ions and thiocyanate ions.

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons [Fe(SCN)]^{2+}(aq)$$
$$\text{pale yellow} \quad \text{Colourless} \qquad \text{deep red}$$
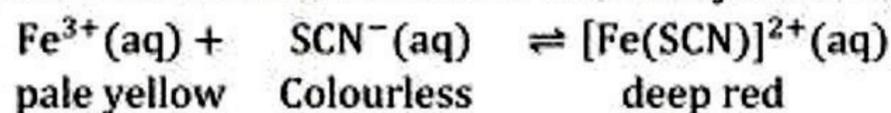
The equilibrium constant for the above reaction can be written as.

$$K = \frac{[Fe(SCN)]^{2+}}{[Fe^{3+}][SCN^-]}$$

Where, $[Fe(SCN)]^{2+}$, $[Fe^{3+}]$ and $[SCN^-]$ are the equilibrium concentrations of the respective species while K is the equilibrium constant. For a particular reaction, the value of K is constant at a particular temperature. When the concentration of any species involved in the equilibrium is disturbed, the concentration quotient,

$$\frac{[Fe(SCN)]^{2+}}{[Fe^{3+}][SCN^-]}$$

remains no longer equal to K. To re-establish the equilibrium, the ions interact in such a way that the concentration quotient again becomes equal to the equilibrium constant K.

A. **Effect of increasing concentration of ferric ions**

When ferric chloride solution is added to the red solution containing ferric ions, thiocyanate ions and ferric-thiocyanate complex, the concentration of ferric ions increases and therefore, the concentration of thiocyanate ions should decrease or that of $[Fe(SCN)]^{2+}$ should increase to keep concentration quotient equal to the equilibrium constant at a given temperature. Therefore, an increase in the concentration of ferric ions results in more thiocyanate ions combining with ferric ions to give more of $[Fe(SCN)]^{2+}$ complex and therefore, the colour intensity of red-solution increases. Thus, an increase in the concentration of Fe^{3+} ions shift the above equilibrium in the forward direction.

B. **Effect of increasing concentration of thiocyanate ions**

Since the thiocyanate ion is in the denominator in the equilibrium law equation, the addition of more and more of thiocyanate results in more ferric ions reacting with thiocyanate ions to give more of $[Fe(SCN)]^{2+}$ complex. Hence, the colour intensity of red-solution increases. Thus, an increase in the concentration of SCN^- ions shift the above equilibrium in the forward direction.

C. **Effect of increasing the concentration of potassium ions**

When potassium chloride is added to the red solution, the concentration of K^+ ions increases. It affects the equilibrium between potassium and thiocyanate ions.

$$KCNS \rightleftharpoons K^+ + SCN^-$$
$$KCl \longrightarrow K^+ + Cl^-$$
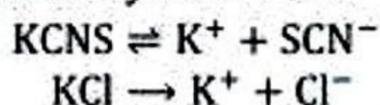
An increase in the concentration of K^+ ions shift the equilibrium in the backward direction. This results in a decrease in the concentration of SCN^- ions which in turn shifts the equilibrium in the backward direction. In other words, some of the $[Fe(SCN)]^{2+}$ complex dissociates to give Fe^{3+} ions and SCN^- ions. Due to a decrease in the concentration of $[Fe(SCN)]^{2+}$ the intensity of the red colour decreases. Thus, an increase in the concentration of K^+ ions shift the above equilibrium in the backward direction.

PROCEDURE

(i)   Take a 250 ml beaker thoroughly washed and clean.

(ii) Put 10 ml of 0.1 M $FeCl_3$ solution in it by using a measuring cylinder.

(iii) Add 10 ml of 0.1 M KSCN solution with the help of a measuring cylinder.

(iv) A deep red colour is obtained due to complex formation $[Fe(SCN)]^{2+}$ (aq).

(v)   Dilute the above deep red solution by adding 50 ml of distilled water.

(vi) Take four test tubes and label them as A, B, C and D. Add 10 ml of the deep red solution to each of the four test-tubes.

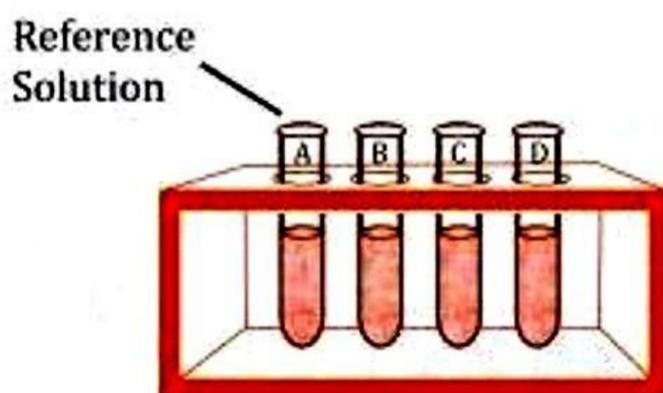(vii) Arrange the test tubes in a test tube stand [Fig. 1].

Reference Solution

Fig' 1 Placing of test tubes.

(viii) Add 5 ml of distilled water to test tube A; 5 ml of 0.1M $FeCl_3$ solution to test tube B; 5 ml of 0.1 M KSCN solution to test tube C and 5 ml of 0.1 M KCl solution to test tube D.

(ix) Shake all the tubes well.

(x)   Now compare the intensity of the colours in test tubes, B, C and D with the red colour in test tube A taken as a reference tube.

(xi) The intensity of the red colour corresponds to the concentration of complex $[Fe(SCN)]^{2+}$ and if the concentration of this ion increases, the colour intensity will also increase.

OBSERVATION

Test tube	A substance added at equilibrium	Change in colour	Effect on the concentration of $[Fe(SCN)]^{2+}$	Shift of equilibrium
A	5 ml of water	Reference colour	—	—
B	5 ml of 0.1 M $FeCl_3$ solution	Colour deepens	Increases	Towards right
C	5 ml of 0.1 M KSCN solution	Colour deepens	Increases	Towards right
D	5 ml of 0.1 M KCl solution	Colour becomes lighter	Decreases	Towards left

CONCLUSION

An increase in the concentration of either of the reactants (Fe^{3+} ions or SCN^- ions) shifts the equilibrium in the forward direction (**towards the right**), on the other hand, a decrease in the concentration of any of the reactants shifts the equilibrium in the backward direction (**towards left).**

PRECAUTIONS

(i) Use tubes of almost identical diameter.

(ii) Dilute solutions of thiocyanate should be used.

(iii) The intensity of the colour of a solution should be compared by keeping it and reference side by side and then observing from the top.

VIVA VOCE

Q 1. Define equilibrium in chemical terms.

Ans. Equilibrium is a state in a chemical reaction where the rates of the forward and reverse reactions are equal, and the concentrations of the reactants and products remain constant over time.

Q 2. How does Le Chatelier's Principle apply to this equilibrium system?

Ans. Le Chatelier's Principle states that if a system at equilibrium is subjected to a change in concentration, temperature, or pressure, the system will adjust to counteract the change and restore equilibrium. In this case, changes in concentration of ferric ions or thiocyanate ions will shift the equilibrium position accordingly.

Q 3. What factors could affect the equilibrium between ferric ions and thiocyanate ions besides concentration?

Ans. Temperature and pressure can also affect the equilibrium. Changes in temperature can shift the equilibrium position, while changes in pressure (for reactions involving gases) can affect the equilibrium only if there is a change in the number of moles of gas.

Q 4. How would increasing the concentration of ferric ions affect the equilibrium position?

Ans. Increasing the concentration of ferric ions would shift the equilibrium position to the right, favouring the formation of more $FeSCN^{2+}$ ions.

Q 5. Describe a possible experimental setup to study the effect of concentration on this equilibrium.

Ans. A spectrophotometer could be used to measure the absorbance of $FeSCN^{2+}$ ions at different concentrations of ferric ions and thiocyanate ions. Solutions of known concentrations could be prepared, and the equilibrium constant (Kc) could be determined from the absorbance values.

Q 6. What is the significance of the equilibrium constant (Kc) in this experiment?

Ans. The equilibrium constant, Kc, indicates the extent of the reaction at equilibrium and helps in predicting the direction in which the reaction will proceed under given conditions.

Q 7. How can the effect of concentration on the equilibrium be quantified?

Ans. By varying the concentrations of ferric ions and thiocyanate ions, plotting a graph of concentration versus absorbance, and analysing the data using the equilibrium constant expression.

Q 8. How might the results of this experiment be applied in real-world scenarios?

Ans. Understanding the effect of concentration on this equilibrium can be applied in industries such as pharmaceuticals and environmental monitoring where precise control of chemical reactions is necessary.