Kevin Zhang Lab Report 6

Introduction

In this lab, we will be observing the common ion effect in action. In particular, we will be introducing additional ions to saturated equilibrium systems, and observing any changes that result.

Chemical Responsibility

Some of the chemicals are irritants or corrosive. In particular, iron (III) chloride is a skin irritant and corrosive, potassium thiocynate and ammonium chloride are toxic when ingested. Silver Nitrate stains skin and is toxic. Ammonia is a respiratory irritant. Gloves and goggles should be worn at all times.

Report Sheet

Adding HCl to saturated NaCl

Observations:

- White / clear particles form along where the HCl makes contact with the NaCl, and they collect along the bottom

Explanation:

- Prior to adding the HCl, the NaCl was dissolved in the water. This was at equilibrium. The addition of chloride caused the formation of table salt (NaCl) particles because there is now too much NaCl to stay dissolved at the current temperature.

Adding Fe^{3+} to [$FeNCS^{2+}$]

Observations:

- The color intensifies, becoming a darker orange.

Explanation:

- The bright red ion is $FeNCS^{2+}$. By increasing the concentration of Fe^{3+}, more of the bright red ion is formed in order to reach equilibrium again.

Adding SCN^- to [$FeNCS^{2+}$]

Observations:

- the color intensifies, becoming a darker orange

Explanation:

- the bright red ion is $FeNCS^{2+}$. By increasing the concentration of SCN^-, more of the bright red ion is formed in order to reach equilibrium again.

Adding Ag^+ to [$FeNCS^{2+}$]

Observations:

- the solution becomes cloudy / milky white
- orange color fades with more drops of silver nitrate

Explanation:

- the silver nitrate is reacting with something else in the solution (likely the Cl^-) and forming a different compound. This disrupts the equilibrium.

Adding NH^+

Observation:

- the pink color fades away and solution looks clear

Explanation:

- the pink color is an indicator of the pH of the solution. Adding more NH^+ increases the conjugate acid concentration, which lowers the pH.

Adding HCl

Observation:

- the pink color fades away and the solution looks clear

Explanation:

- HCl is an acid, and the pink color is a pH indicator for basic solutions. Adding more HCl lowers the pH, which causes the solution to go clear.

Sample Calculations

None

Discussion of Results

The results show Le Chatelier's Principle in action. Several of the solutions produced a visible result when the equilibrium of a solution was disturbed, with the change of concentration of ions causing a shift to either the product or reactant side (depending on the reaction).

Post-Lab Questions

1. Define Le Chatelier's Principle Le Chatelier's Principle states that if a chemical system in equilibrium is disrupted (moved away from equilibrium), the chemical system will work towards equilibrium again. This can be done by either converting products to reactants, or vice versa.
2. Why was it necessary to dilute the $Fe(NCS)^{2+}$ solution? Because the solution was already pretty orange/red, which would make it difficult to detect color changes.
3. Consider the following reaction:

$$[CoCl_4]^{2-}(\text{aq, blue}) + 6H_2O(\text{l}) \rightleftharpoons [Co(H_2O)_6]^{+}(\text{aq, pink}) + 4Cl^{-}(\text{aq})$$

   1. Solution turns blue when placed in hot water bath. Is above reaction exothermic or endothermic? The reaction is exothermic.
   2. Explain in terms on Le Chatelier's Principle The reaction must be exothermic, because heat is considered a product in the above reaction. When excess product was added to the system via the hot water bath, the reaction tried to return to equilibrium by producing more reactant. This turned the solution blue.
4. Milk of magnesia dissolves in water to produce a cloudy solution. The equilibrium reaction is:

$$Mg(OH)_2(\text{s, white}) \rightleftharpoons Mg^{2+}(\text{aq}) + 2OH^{-}(\text{aq})$$

   1. What observations would you see if a small amount of concentrated HCl was added to equilibrium mixture? The solution would become more clear, and less cloudy.
   2. Write the net ionic equation for the reaction in part A.

$$H^{+}(\text{aq}) + OH^{-}(\text{aq}) \rightarrow H_2O(\text{l})$$

   3. Explain your observations in terms of Le Chatelier's Principle. The HCl reacts with the OH^{-}, reducing the products in the original equilibrium. This causes the reaction to consume more reactants to produce more product.

Conclusion

This lab is an analysis of the Le Chatelier's Principle in action. We observe it in multiple reactions, on both the product and reactant side.