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Chemical Equilibrium

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ABSTRACT

Equilibrium chemistry occurs when the concentrations of reactants and products remain constant over time. If an external change affects the system, the reaction shifts to establish a new equilibrium, according to Le Chatelier's Principle. This experiment tested equilibrium involving iodine dissolved in two immiscible solvents: organic CCl4 and water. The aim was to identify factors affecting chemical equilibrium and to determine the equilibrium constants for the reaction I2 + I− ↔ I3−. Equilibrium is influenced by temperature, pressure, volume, and concentration. The reaction is reversible, meaning products can revert to reactants until equilibrium is reached. In this practicum, concentration and volume were observed. The method involved saturating I2 in CCl4 in two different Erlenmeyer flasks, then titrating the later solution with thiosulfate to measure I2 content. Titration was done quickly after adding starch indicator due to the importance of equilibrium time. Two experiments were conducted: the first to determine the distribution constant (KD) by measuring I2 in both CCl4 and water, and the second to find the equilibrium constant (KC) by adding KI solution to I2 in CCl4. The obtained values were KC = 66.176 M and KD = 32.

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1. Introduction

Equilibrium chemistry is a dynamic process. Reactions occur at the same rate but in opposite directions at the reaction front and the reaction back. Bonding occurs when the equilibrium of a reaction is broken or formed between reactant and product molecules. At the beginning of the reaction, the concentration of reactants is high. Then, collisions between molecules form product molecules. Concentration of product molecules occurs when enough reactions take place; then, the opposite

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(formation of "reactants" from "products") begins to take place [1]. Reaction chemistry generally takes place reversibly, which is also called two-way, and involves lots of walking. It is not perfect, which means that the only reaction is to walk, reach the point, and finally stop, leaving behind substances that do not react. At a certain temperature, pressure, and concentration, the period at which the reaction stops is consistent, and the relationship between the reactant and product concentrations still exists. The speed of the reaction at equilibrium moves to the right at the same rate as the speed of the reaction to the left [2].

Solubility Iodine (I_2) is highly soluble in water and insoluble in carbon tetrachloride (CCl_4). The second solvent, however, cannot mix with water and forms two layers: water on top and CCl_4 on the bottom. When a solution of I_2 in CCl_4 is added to water and shaken, the second solvent will separate and return at the end. Then, the second solvent separates and returns. As a result, I_2 is distributed into two solvents in proportion to the two solvents. At a constant temperature, the sum of the concentration ratios of the two solvents remains constant, a phenomenon known as constant coefficient distribution or partition coefficient. The KD value can be determined by measuring the amount of I_2 in the two solvents. Although I_2 dissolves poorly in water, it can grow and multiply if the water contains potassium iodide (KI) due to the formation of triiodide (I_3^-) complex ions. The reaction formation is back-and-forth until equilibrium is finally formed with constant equilibrium (Kc). The Kc value is very large. This means that the equilibrium is highly skewed to the right, so $[I_3^-] >> [I_2]$. This equilibrium can be achieved by adding a KI solution to an I_2 solution in CCl_4. After shaking and leaving it to cool, the solvent separates back out, resulting in a higher concentration of I_2 in the water compared to the CCl_4 [3].

An equilibrium reaction is closely tied to Le Châtelier's principle. A system in equilibrium tends to maintain its balance. If there is an outside influence, the system will change so that balanced circumstances quickly return. Henri Le Châtelier, a French chemist, discovered that if a chemical reaction in equilibrium undergoes a state change (i.e., accepts external action), the reaction will reach a new equilibrium with a shift to overcome the changes (i.e., commit a reaction in response to changing circumstances). This is called Le Châtelier's principle. The "[ ]" sign represents equilibrium concentration. The reaction rate of a chemical reaction at a constant temperature is proportional to the product of the concentrations of the reacting substances. Reaction chemistry moves to an equilibrium dynamic where there are reactants and products. However, position No. 2 again has a tendency to change. Sometimes, the concentration of the product is much greater than the concentration of unreacted reactants in mixture equilibrium. In this case, the reaction is said to be "perfect." In equilibrium, catalysts increase the speed of the reaction, and the reaction returns to the same state. Catalyst: There is no change in the amount relative to the balance. Although equilibrium means no change, a catalyst can alter the time required to reach equilibrium. A reaction may proceed at the appropriate rate only at very high temperatures, but it can occur quickly at far lower temperatures if a catalyst is used [5]. A catalyst in equilibrium can speed up the reaction rate so that equilibrium is quickly achieved. Catalysts are substances that can increase the rate of reaction until equilibrium is reached, at which point the reaction becomes permanent. Equilibrium is reached when the reaction occurs permanently. For example, a catalyst in cell material burns hydrogen, breaking the molecule of oxygen (cathode) into oxygen atoms or ions, which react with hydrogen atoms or ions from the anode [6]. Equilibrium chemistry is a dynamic process where the reaction progresses and comes back at the same rate but in opposite directions. Bonding to reaction directions: Reactions in equilibrium will break or form between the reactant and product molecules. When you concentrate the beginning reactant, the beginning reactant increases, so collisions between the molecules form a molecule, a product molecule. When you concentrate the product, the product becomes big enough, and the reaction reverses (the formation of "reactants" from "products" and "products" from "reactants" begins taking place). Reaction chemistry generally takes place reversibly and can be called a two-way reaction. It also involves lots of walking, which means reactions are only perfect reactions. Perfect reactions are reactions that stop at a certain point, leaving substances that do not react. At a certain temperature, pressure, and concentration, the reaction will stop. The relationship between the concentration of reactants and products still exists at the point at which the reaction stops. The speed of the reaction at equilibrium moves to the right at the same rate as the speed of the reaction to the left [7].

2. Materials and Method

This experiment required Erlenmeyer flasks, glass measuring pipettes, measuring pipettes, burettes, and spray bottles. The materials used were an I_2 solution saturated in CCl_4, 0.1 M KI, 0.02 M $Na_2S_2O_8$, a 1% indicator solution, and solid KI crystals.

Provide two Erlenmeyer flasks, labeled A and B. Add 20 mL of solution to Erlenmeyer A and add I_2 to CCl_4. For the procedure to determine KD, add 200 mL of distilled water to Erlenmeyer A. Close the Erlenmeyer tightly and shake it vigorously. Then, place it in a thermostat at 30°C for 30-60 minutes. Occasionally, remove the Erlenmeyer and shake it. Once equilibrium is achieved, remove 5 mL of solution from the CCl_4 layer (below the water layer) using a pipette. Add two grams of solid KI crystals and 20 mL of water, and then homogenize. Perform the titration with a standard sodium thiosulfate solution. Add starch (10 mL) to the pale yellow solution. Record the volume of thiosulfate used (V1). Take 50 mL of the water layer solution and titrate it with the thiosulfate solution. Record the volume of thiosulfate used (V_2). Then, determine the KC value. For the procedure to determine the KC value, 200 mL of solution was added to standard KI 0.1 M in Erlenmeyer B. The Erlenmeyer was closed tightly and shaken vigorously. Afterwards, it is placed in a thermostat at 30°C for 30-60 minutes. Occasionally, remove the Erlenmeyer flask and shake it. Once equilibrium is achieved, use a pipette to take 5 mL of solution from the CCl_4 layer (below the water layer). Add two grams of solid KI crystals and 20 mL of water, and then homogenize. Perform the titration with the standard sodium thiosulfate solution. Add starch (10 mL) to the paleyellow solution. Record the volume of thiosulfate used (V_3). Take 50 mL of the water layer solution and titrate it with the thiosulfate solution. Record the volume of thiosulfate used (V_4). Then, determine the KC value.

3. Result

Erlenmeyer A

In layer I2 in CCl4

[2] CCl 4 = 0.002 M

In the water layer

[2] water = 0.000075 M

Coefficient Distribution (KD) = 2

Erlenmeyer B

[2] CCl 4 = 0.0006 M

[I2] water = 0,001235 M

[I2] free = 18,775 x 10 $*$ −6 M

[I3-] =0,001216 M

[I-] = 0,098 M

KC = 66,175 M

4. Discussion

In this practicum, we tested equilibrium chemistry. Equilibrium chemistry occurs when the rate of reaction is the same from reactant to product. The rate of reaction is the speed of the product.

In this experiment, we used a mixture of solvents that don't mix with the organic solvent CCl_4. The concentration of the organic solvent CCl_4 is determined by the concentration of the heterogeneous iodine in equilibrium with the heterogeneous iodine in two immiscible solvents. Determination of the CCl_4 concentration and the I_2/I_3 balance in an aqueous solution is carried out by balancing the KI solution with CCl_4 and the I_2 solution in CCl_4. After achieving equilibrium, the second solution was separated, and each solution was titrated with sodium thiosulfate to determine the I_2 content. Mixing an I_2 solution in CCl_4 with a KI solution produces a purple solution.

After a few silent moments, the solution apparently separated. The yellow section is concentrated, while the lower part is an old purple color. The yellow part is iodine dissolved in water. Based on observation, the separation solution is possibly in equilibrium. The iodine distributed in the CCl_4 solution has been achieved.

To determine constant equilibrium, first count the total iodine concentration, then the iodine concentrations in the CCl_4 and water solutions, as well as the I_3^- and I^- levels. Once the concentrations of all existing species are balanced in known circumstances, the mark of constant equilibrium can be determined. Equilibrium occurs when iodine dissolved in water as potassium iodide undergoes the following reaction: $I_2 + I^- \leftrightarrow I_3$. This reaction moves to dynamic equilibrium, where there are reactants and products, but the position is not fixed and has a tendency to change. Sometimes, the concentration of the product is much greater than the concentration of unreacted reactants in the equilibrium mixture, so the reaction is said to be "perfect" [8].

After that, the solution was left undisturbed until a clear difference was seen in the second layer. Then, indicator starch was added. This indicator serves a purpose. For endpoint titration, changes in color indicate the endpoint. This indicator binds the released I_2 from its bond with water or CCl_4. The entry of I_2 into starch produces an old blue color in the titrated solution.

Next, the second layer solution was titrated with 0.02 M $Na_2S_2O_8$. Titration was done until the solution changed from a purple-blue color to clear. Equilibrium time is very important, so titration should be done as soon as possible. It is possible to add the starch indicator after the solution is prepared. After performing the calculations, the KC value was found to be 66.176, and the KD value was found to be 32.

5. Conclusions

This experiment tested equilibrium chemistry, focusing on the equilibrium that occurs when the reaction rate from reactant to product is equal to the reaction rate from product to reactant. The experiment also involved the law of distribution, in which iodine was dissolved in two immiscible solvents: CCl_4 and water. To determine the concentrations of species in equilibrium, a heterogeneous equilibrium of iodine in the two solvents was established. The concentrations of I_2 and I_3^- in the aqueous solution were determined by titration with sodium thiosulfate. Mixing an I_2 solution in CCl_4 with a KI solution resulted in a purple solution. After settling, it was observed that the upper layer was yellow, indicating the presence of iodine in water, and the lower layer was purple, indicating the presence of iodine in CCl_4. This suggests that equilibrium was achieved with the iodine distributed between the two solvents. Then, the equilibrium constant (KC) and distribution constant (KD) were determined based on the concentrations of the various species. Adding a starch indicator facilitated endpoint titration with 0.02 M

$Na_2S_2O_8$ until the solution changed from purple-blue to clear. The obtained KC value was 66.176, and the obtained KD value was 32.

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