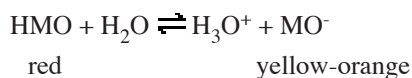


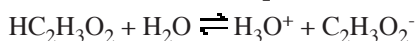
phenolphthalein to the other. Add one drop of 6 M HCl to each test tube. Add two drops of 6 M NaOH to each test tube. Record your observations and explain what happened using Le Châtelier's principle. Here is the generic indicator equation rewritten for methyl orange:



HCl increases the H_3O^+ , and NaOH decreases the H_3O^+ . Rewrite the generic equation in a similar way for phenolphthalein in the explanation section on the next page.

Dispose of the solutions in the sink.

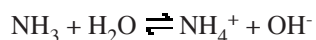
Part 3: Acetic Acid/Acetate Equilibrium.



Pour about 1 milliliter of 0.1 M $\text{HC}_2\text{H}_3\text{O}_2$ into each of two test tubes. Add a drop of methyl orange solution to each test tube. Add 1 M $\text{NaC}_2\text{H}_3\text{O}_2$ solution to one test tube a drop at a time until there is a difference in color between the two solutions. Record your observations and explain what happened using Le Châtelier's principle. When sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$, dissolves in water, it breaks up into the Na^+ and the $\text{C}_2\text{H}_3\text{O}_2^-$ ions. The methyl orange changes color with changing H_3O^+ concentration, so refer to Part 2 of the experiment to help you interpret what the addition of acetate ions did to the equilibrium.

Dispose of the solutions in the sink.

Part 4: Ammonia/Ammonium Equilibrium.



The OH^- ion concentration controls the H_3O^+ ion concentration in water solutions. As the OH^- ion concentration gets larger, the H_3O^+ ion concentration gets smaller, and vice versa. The $[\text{OH}^-]$ in the original aqueous solution of ammonia, NH_3 , is about 10^{-3} , which makes the $[\text{H}_3\text{O}^+]$ about 10^{-11} .

Pour about 1 milliliter of 0.1 M NH_3 into each of two test tubes. Add a drop of phenolphthalein solution to each test tube. Add 1 M ammonium chloride, NH_4Cl , solution to one test tube a drop at a time until there is a difference in color between the two solutions. Record your observations and explain what happened using Le Châtelier's principle. When NH_4Cl dissolves in water, it breaks up into the NH_4^+ and the Cl^- ions. As you were instructed in Part 3, use the information from Part 2 to help you interpret what the addition of ammonium ions did to the equilibrium.

Dispose of the solutions in the sink.

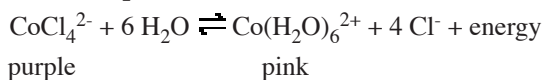
Part 5: Iron/Thiocyanate Equilibrium.



Pour about 1 milliliter of deionized water to each of three test tubes. Add 1 drop of 0.1 M $\text{Fe}(\text{NO}_3)_3$ and 1 drop of 0.1 M KSCN to each test tube. Add two more drops of 0.1 M $\text{Fe}(\text{NO}_3)_3$ to one of the tubes, and two more drops of 0.1 M KSCN to another of the test tubes. Try to compare the colors of the solutions as quantitatively as possible. Record your observations and explain what happened using Le Châtelier's principle.

Dispose of the solutions in the sink.

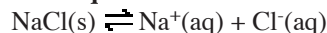
Part 6: Temperature Effect.



Pour about 2 milliliters of 0.15 M CoCl_2 (in methanol) into a test tube. Add deionized water drop by drop with stirring until the blue color *just changes* to pink. Do not add excessive water. Place the test tube in a water bath in the fume hood. The temperature in the water bath should be 65 to 70 °C. When the color changes, remove the test tube from the hot water bath, and cool it by swirling it in tap water. (If you do not see a color change, you may have added too much water. Remake the solution. Take care to follow directions.) Record your observations and explain what happened using Le Châtelier's principle.

Dispose of the solution in the waste container marked "Cobalt Waste".

Part 7: Saturated Equilibrium.



Pour about 1 milliliter of saturated NaCl solution into a test tube. Add 12 M HCl drop by drop with stirring until you see a change taking place in the test tube. Continue to add 3 or 4 more drops, pausing after each drop. Record your observations and explain what happened using Le Châtelier's principle.

Dispose of the solution in the sink.

OBSERVATION TABLE

Rewrite the chemical equation for each part in the explanation .

Part 1 Observations

Explanation

Part 2 Observations

Explanation

Part 3 Observations

Explanation

Part 4 Observations

Explanation

Part 5 Observations

Explanation

Part 6 Observations

Explanation

Part 7 Observations

Explanation

QUESTIONS

1. What you did in Parts 3, 4, 5, and 7 are examples of the *common ion effect*. Why do you suppose it has that name? Give an explanation of the common ion effect based on what you did and what you observed.
2. Rewrite the chemical equation from Part 6 using ΔH notation for energy. Simply indicate if the ΔH is + or -.
3. Write the following equation in the manner the equation in Part 6 is written (note the energy term):
 $2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{H}_2\text{O}_2(\text{aq}) \quad \Delta H = 196.1 \text{ kJ}$
4. In Part 7, the $[\text{Cl}^-]$ in saturated NaCl is 5.4 M at room temperature. Assume that you had 1.00 ml of the saturated solution, and that you added 0.50 ml of 12 M HCl. What is the $[\text{Cl}^-]$ after you added the HCl. (When two solutions contain the same component, the numerator consists of the sum of the volume times the concentration for each solution. The denominator is the total volume.
5. Ammonia, NH_3 , has a distinctive smell. In Part 4, how do you think the smell of the solution would be affected by the addition of NaOH? Explain.
6. Write out the K_{eq} expression for each reaction shown in the seven parts of the experiment. For reactions in water solutions, $[\text{H}_2\text{O}]$ is not shown in the expression because it is usually a large and relatively constant value.

1**2****3****4****5****6****7**