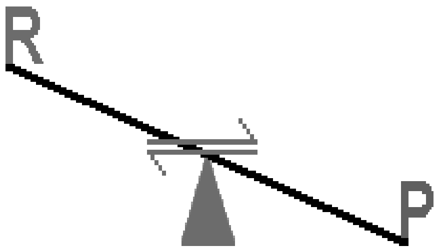
4—Le Chatelier's Principle



Name: ______________________________

Date: ______________________________

Section: ______________________________

Objectives

- To Investigate Le Chatelier's Principle
- To relate macroscopic observations to microscopic changes
- To review writing net ionic equations

Pre-Laboratory Requirements

- Read Chapter 17.6 in Silberberg
- Review writing net ionic equations (Silberberg 4.2)
- Pre-Lab Questions (if required by your instructor)
- Laboratory Notebook—prepared before lab (if required by your instructor)

Safety Notes

- Eye protection must be worn at all times

Discussion

Le Chatelier's principle states that if a system in equilibrium is subjected to an external stress (i.e., change in concentration, pressure, volume, or temperature), the system adjusts itself in such a way that it minimizes the effect of this stress. It is used to predict how the equilibrium concentrations will change as the system adjusts to re-establish the equilibrium. As you know from the previous experiment, chemical equilibrium represents a balance between forward and reverse reactions to give a fixed value for the equilibrium constant (K) at a fixed temperature. Changes in experimental conditions can disturb this balance. Since K has fixed value, the concentrations will change in a way to produce equilibrium concentrations that give the same value for K. Think about the thermostat in your house that is set to keep the room at a comfortable temperature (its equilibrium). As cold (or hot) air leaks into the house, the thermostat will turn the heater on (or off) in order to bring the room back to the set temperature (its equilibrium condition).

To understand what is meant by a stress, consider the following simple reaction at equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2\,NH_3(g) \qquad \Delta H = -92.6\ \text{kJ}$$

One way to stress this system is to add a reactant (N_2 and/or H_2). The system is no longer at equilibrium and will adjust to return to equilibrium. Some reflection revels that there is now an excess of reactant. The system will need to use more reactants and produce more products to return to equilibrium. A similar conclusion would be reached if NH_3 were removed. In this case the stress is relieved by forming more NH_3 so the forward reaction is favored (the equilibrium will shift to the right). Conversely, if more NH_3 were added, the reverse reaction would be favored and the equilibrium will shift to the left producing more N_2 and H_2.

Since the equilibrium constant is independent of the pressure, changing the pressure causes no shifts in the equilibrium concentrations if the volume does not change. Since [concentration] = n/ V, changing the volume will stress the system. For example, increasing the volume will decrease the concentrations of all species and the reaction will no longer be in equilibrium. Since there are 4 moles of gas on the left side (reactants) and 2 moles of has on the right side (products) of this equation, this change will produce a reaction quotient that is too large. To re-establish equilibrium, the concentration of products must decrease meaning more reactants must form (the equilibrium shifts to the left). In general, increasing the volume will cause the equilibrium to shift in the direction that produces the largest number of moles of gas. Decreasing the volume shifts the equilibrium toward the smaller number of moles of gas. Volume changes have little effect when all species are in solution since the volume of the solution is independent of the volume of the container. .

Changing the temperature changes the value of the equilibrium constant. These changes can be rationalized by considering heat to be a product for an exothermic reaction or a reactant for an endothermic reaction. NOTE: all equilibrium reactions have an exothermic and an endothermic reaction occurring simultaneously. Adding heat favors the endothermic process since there is more heat to react and the equilibrium will shift in that direction. For the reaction above, heating the system will shift the reaction to the left to consume the added energy (since that is the exothermic reaction). Cooling the system (removing heat) will produce the opposite change.

In this experiment, you will prepare an equilibrium system for the following reaction

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

and study the response of this system to various stresses. The results and observations are to be explained in terms of Le Chatelier’s principle. Net ionic equations for all of the reactions occurring at equilibrium will be needed to justify your conclusions.

Procedure

1. Add 125 mL of DI water to a clean 400 ml beaker
2. Add 10 drops of 1 M KSCN and 10 drops of 1 M $Fe(NO_3)_3$ solution to the DI water (*Note: This will be your stock solution and will have an intense red color due to the formation of the complex* $FeSCN^{2+}$)
3. Obtain 9 clean medium-sized test tubes and label them 1 through 9
4. Add approximately 10 mL of the stock solution into each test tube
5. Test tube 1 will be kept as a control and the following reagents will be added into test tubes 2-8, respectively:

Test tube #	Treatment
1.	Only stock solution
2.	2 mL of 1 M $Fe(NO_3)_3$
3.	1 mL of 1 M KSCN
4.	0.5 mL (10 drops) of 0.1 M $AgNO_3$
5.	2 mL of concentrated HCl
6.	1 mL of 1 M Na_3PO_4
7.	1 mL of 0.1 M $Na_2C_2O_4$
8.	Several crystals of solid NaF
9.	Only stock solution

6. Cover each test tube and mix.
7. Heat Test tube 9 (*which only has the stock solution in it*) in a hot water bath
8. Compare the color intensity of each of the test tubes with that of the control test tube number 1. In each test, record your observations and explain the results in terms of Le Chatelier's principle
9. Include net ionic equations in your explanations.

Note: All of the color comes from $FeSCN^{2+}$***. The darker the color, the larger the concentration of*** $FeSCN^{2+}$ ***in the tube.***

Use the complexes and precipitates in **Table 1** to explain your observations and write the equations of any reactions that occur:

Table 1. Complexes and Precipitates

	Ag^+	Cl^-	PO_4^{3-}	F^-	$C_2O_4^{2-}$
Fe^{3+}:		$FeCl_4^-$	$FePO_4$(s)	FeF_6^{3-}	$Fe(C_2O_4)_3^{3-}$
SCN^-:	AgSCN(s)				

Data:

Test Tube #	Reagent Added	Observation	Explanation
1			
2			
3			
4			
5			
6			
7			
8			
9			