**Le Châtelier’s Principle – Supplemental Worksheet**

Determine how the reaction will shift to re-establish a new equilibrium following a change. Will the reaction shift to the left (toward the reactants), to the right (toward products), or would there be no change?

A. \[ 2\text{SO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2\text{SO}_3 (g) \] (This process is **EXOTHERMIC**)

1. If we add more \( \text{SO}_2 \) gas to the reaction chamber. **To the left.** The reaction would drive backwards to re-establish equilibrium (more reactants made).

2. If we decrease the volume of the reaction chamber. **To the right.** This reaction deals with gases! So decreasing the volume changes the pressure (raises the pressure). We have more moles of gas on the reactants side (3 moles) than on the products side (2 moles). A change in pressure will be cubed for the reactants but only squared for the products. The reaction forwards towards to products to relieve some of the pressure off of the reactants.

3. If we lower the temperature by removing heat from the reaction chamber. **To the right.** The process is exothermic, and we could think of heat as a product. The true reason the reaction shifts is because the equilibrium condition actually changes. The value of \( K \) increases with decreasing \( T \) in this exothermic case. The result is a shift to reactants because the reactant/product concentrations change to fit this new equilibrium condition (increasing \( K \) means more products), not that the heat is actually a product. Either way you think about it, removing the heat product will drive the reaction forward to the right.

B. \[ \text{PbCl}_2 (s) \rightleftharpoons \text{Pb}^{2+} (aq) + 2\text{Cl}^- (aq) \] (This process is **ENDOTHERMIC**)

1. If we add \( \text{NaCl} \) to the solution. **To the left.** This is due to the common ion effect as applied to the Le Châtelier Principle. The concentration of the \( \text{Cl}^- \) ion product would increase, so the reaction would shift towards reactants. We could use \( Q \) and \( K \) as well: at the initial moment of addition \( Q \) would be larger than \( K \). This means the reaction would have to shift back towards the reactants to reestablish equilibrium.

2. If we raise the temperature by adding heat to the solution. **To the right.** Because the reaction is endothermic, energy must be added to the system for the reaction to occur. Therefore, we could consider heat to be a reactant. The true reason the reaction shifts is because the equilibrium condition actually changes. The value of \( K \) increases with increasing \( T \) in this endothermic case. The result is a shift to products.
because the reactant/product concentrations change to fit this new equilibrium condition (increasing K means more products), not that the heat is actually a reactant. Either way you think about it, raising T would shift the reaction toward the products (toward the right).

3. If we add more solid PbCl\textsubscript{2}. No change. The concentrations of the ions are already at equilibrium, so no more solid needs to dissolve. Thus, adding more solid will not shift the reaction. This is not to say that NONE of the added solid would dissolve, but rather there is a dynamic equilibrium between the rate of the solid dissolving into ions and the rate the ions precipitating out of solution and back onto the solid. Adding more solid would not disrupt this rate because the solution was already at equilibrium.

4. If we add pure water to dilute the solution. To the right. The dilution lowers the concentration of both of the ions in solution (more volume means lower concentration). So in order to raise the concentrations back up to their equilibrium concentrations more solid reactants would have to dissolve off into ion products (shift to the right).

C. \[ A_x \ (g) \ + \ \text{heat} \leftrightarrow B_y \ (g) \ + \ C_z \ (g) \]

1. If inert argon gas is added to the reaction chamber until the total pressure doubles. No change! The addition of an inert gas lowers the partial pressures of the reactants and the products, but the change in the partial pressures occurs without disturbing equilibrium ratios of the partial pressures.

2. If heat is removed from the reaction chamber. To the left. It can be inferred from the equation that the reaction is endothermic, so removing heat is “like” removing a reactant. The true reason the reaction shifts is because the equilibrium condition actually changes. The value of K decreases with decreasing T in this endothermic case. The result is a shift to reactants because the reactant/product concentrations change to fit this new equilibrium condition (decreasing K means more reactants), not that the heat is truly a reactant.

3. If A\textsubscript{x} is added to the reaction chamber. To the right. A\textsubscript{x} is a reactant, so adding more reactant will shift the equilibrium toward the products (to the right).

D. \[ 6\text{CO}_2 \ (g) + 6\text{H}_2\text{O} \ (l) + \text{light/heat} \leftrightarrow \text{C}_6\text{H}_12\text{O}_6 \ (s) + 6\text{O}_2 \ (g) \]
1. If more oxygen gas is added to the reaction chamber. To the left. Oxygen gas, O₂, is a product, so an increase in products moves the reaction towards the reactants.

2. If we remove CO₂ gas from the reaction chamber. To the left. Carbon dioxide is a reactant and its removal will cause the reaction to shift to create some of the “lost” reactant.

3. If we expand the volume of the reaction chamber. No shift! The expansion will increase the volume, which will decrease the pressure. However, since there is an equal number of gas moles on each side of the reaction (6 moles on each), there is no shift.

4. If we add more H₂O to the reaction chamber. No shift! A change in the amount of liquid water does not influence the equilibrium constant, K, because, by convention, its activity is “1”.

5. Bonus: What is the name of this important reaction? This reaction is the extremely important photosynthesis reaction.