Le Chatelier's Principle Exercises

"If a stress is applied to a system at chemical equilibrium, the equilibrium will shift in such a manner as to counteract the effects of that stress."

This only applies to systems at equilibrium. Other reactions can go to completion:

Many other chemical reactions can only run in one direction, going only from the reactants on the left side of the arrow to the products on the right side of the arrow. These reactions are called "not reversible."

A good example of this might be burning some paper:

$$\textbf{Cellulose(s)} + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

The reaction proceeds until all of either one of the reactants is used up and then it stops. You cannot make the reaction run in reverse. This is usually because of the complexity of one or more of the reactants. For example, imagine putting some carbon dioxide and water together in a beaker and trying to get starch or sugar or any number of other CHO (complex carbohydrate) compounds. It just does not happen!! Typically, reversible reactions are simple one-step reactions in both directions. The making of cellulose by a plant requires many steps, some with different requirements of temperature or time, whereas to break cellulose down to CO_2 and H_2O takes only one step.

1. Assume that the following reaction is in chemical equilibrium:

$$N_2(g) + 3\,H_2(g) \rightleftarrows 2\,NH_3(g) + \textbf{heat}$$

Explain the effect of each of the following changes upon the system in terms of Le Chatelier's Principle and a shift toward either the product or reactant side.

1. More hydrogen is added to the system.
2. Ammonia is removed from the system.
3. Nitrogen is removed from the system.
4. The temperature is raised.
5. The pressure of the system is decreased by doubling the total volume.

2. Which way will the equilibrium shift if the system temperature goes up (heat is added)? Why?

$$2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g) + \text{heat}$$

3. The container holding the following reaction (already at equilibrium) has its volume suddenly reduced by half. Which way will the equilibrium shift to compensate? Why?

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$$

4. The container holding the following reaction (already at equilibrium) has its volume suddenly increased. Which way will the equilibrium shift to compensate? Why?

$$H_2(g) + Cl_2(g) \leftrightarrow 2\ HCl(g)$$

5. The system below is already at equilibrium when some neon is added to the system. What happens to the position of the equilibrium? Does it shift right, left, or no change?

$$H_2(g) + Cl_2(g) \leftrightarrow 2\ HCl(g)$$

6. The system below is already at equilibrium when a catalyst is added to the system. What happens to the position of the equilibrium? Does it shift right, left, or no change? Why?

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$$

The Effect of Temperature on Equilibrium Position

$$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g) + \text{heat energy}$$

a) What effect would increasing the temperature have on the equilibrium position? Explain.

b) What effect would increasing the temperature have on the equilibrium constant? Explain.

Practice writing Keq (equilibrium constants)

$$Keq = \frac{[P]^x}{[R]^y}$$ With products and reactants raised to their coefficients.

n.b. solids and liquids are NOT included as you cannot change the concentration of a solid or a liquid

Write an equilibrium expression for each of the following reactions.

1) $3\ O_2(g) \leftrightarrow 2\ O_3(g)$

2) $N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g)$

3) $NaCl(s) \leftrightarrow Na^+(aq) + Cl^-(aq)$

4) $Cu(s) + 2\ AgNO_3(aq) \leftrightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$

5) $2\ NaBr(aq) + I_2(s) \leftrightarrow Br_2(l) + 2\ NaI(aq)$

Solving for Keq:

Something of a side issue is what are the units of the K_{eq}? For reasons beyond the scope of this lesson, the answer is none. The equilibrium constant does not have any units, it is just a pure number.

1) Calculate the equilibrium constant (K_{eq}) for the following reaction:

$$\mathbf{H_2(g) + I_2(g) <==> 2\ HI(g)}$$

when the equilibrium concentrations at 25 °C were found to be:

$[H_2] = 0.0505$ M
$[I_2] = 0.0498$ M
$[HI] = 0.389$ M

So genius, how do you solve the problem?

2) The same reaction as above was studied at a slightly different temperature and the following equilibrium concentrations were determined:

$[H_2] = 0.00560$ M
$[I_2] = 0.000590$ M
$[HI] = 0.0127$ M

From the data, calculate the equilibrium constant.

Time for a small lecture: Please be very careful in using your calculator to solve these problems. When I solved this problem while writing this tutorial (on December 28, 1998), I first got some really weird looking answer that didn't feel right, so I did it again. Sure enough, I have made an entry error somewhere in the problem. Underscoring my plea for carefulness is the difference between you and me in problem solving. The above problem is routine for me and solely on the basis of experience did I reject my first answer as being wrong (it "felt" wrong). You guys don't have that experience, so you don't have the feel. Yet!!

So, BE CAREFUL.

End of lecture (taken from Chemteam website:
http://dbhs.wvusd.k12.ca.us/webdocs/ChemTeamIndex.html)