Le Chatelier's Principle Worksheet

(print before lab—no additional prelab assignment)

Name ____________________________ Date ______________ Lab section __________________

Iron(III) Ion with Thiocyanate Ion

Chemical equation for the equilibrium:

Mass action expression:

Color of $Fe(NO_3)_3$ solution:

Color of KSCN solution:

Color of reaction mixture (original system):

List your observations and fully explain them using Le Châtelier's Principle for the following:
Effect of adding thiocyanate:

Effect of adding Fe^{3+}:

Effect of adding hydroxide:

Mercury(II) ion combines with thiocyanate to form the compound $Hg(SCN)_2$, which is soluble but undissociated in water, predict the effect upon the equilibrium of your chemical equation above if mercury(II) ion were added to the system.

Ammonium Nitrate with Water

Chemical equation for the equilibrium:

Mass action expression:

Appearance of NH_4NO_3:

Temperature change (hot/cold) for reaction:

Is the reaction endo-or exothermic? ____________________
Rewrite your chemical equation including heat as a reactant or product given your observations:

Predict the effect of raising the temperature:

What happened when you heated the reaction mixture?

Cobalt(II) Ion, Water, and Chloride Ion
Chemical equation for the equilibrium:

Mass action expression:

Concentration effects:
Color of $Co(NO_3)_2$ solution:

Color of HCl solution:

Color of reaction mixture:

List your observations and fully explain them using Le Châtelier's Principle for the following:
Effect of adding HCl:

Effect of adding water:

Effect of adding $CaCl_2$: (the chloride ion in the calcium chloride is not of significance)

Temperature effects

Color of mixture at room temperature:

Color of mixture in boiling water bath:

Color of mixture in ice water bath:

Effect of switching the hot and cold test tubes:

Which of the cobalt complexes predominates in the boiling water bath?

Which of the cobalt complexes predominates in the ice water bath?

Write the equation for the chemical equilibrium including heat as a product or reactant given your observations:

Explain the effect of increasing the temperature on this reaction mixture using Le Châtelier's Principle:

Using your observations from the reactions that you studied today, write a general conclusive statement about how the system copes with the following stresses being applied: added reactant, removing product, increased temperature, and concentrating a reaction mixture (By ”concentrating”, we me removing water, as in the Co experiment. Be sure to consider the mass action expression before answering!). Use specific observations from your experiment to support your claims. How can you predict the shift in the equilibrium position for the reaction for a given “stress”?

Added reactant:

Removing product:

Increased temperature:

Concentrating a reaction mixture:

Chemical Equilibrium and Le Châtelier's Principle

So far this semester, we have made an important assumption in the chemical reactions you have followed in your experiments. The alum, aspirin, and nickel complex syntheses were all considered to proceed to *completion* (this is how you were able to calculate theoretical and percent yield). However, on the molecular level, chemical reactions may (in principle) be considered to be *reversible*. For example, in the following famous reaction:

$$aA + bB \rightleftharpoons cC + dD \qquad (1)$$

where A, B, C, and D are chemical species and a, b, c, and d represent the coefficients of each species in the balanced equation. The double arrow in the equation signifies that the reaction is reversible, meaning that both the forward and reverse reactions are taking place to some extent. Once the *concentrations* of the reactants and the products are no longer changing, the reaction is said to have reached **chemical equilibrium**.

The concentrations of the reactants and products of a reaction at equilibrium can be expressed in the following equation (called the mass-action expression):

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \qquad (2)$$

where K_c is the equilibrium constant and the molar concentrations of reactants and products are given in square brackets raised to the power of the stoichiometric coefficient. Once a reaction has reached equilibrium, the numerical value of the mass-action expression will always equal the constant value K_c as long as the temperature remains constant. If the concentration of any of the species in the reaction changes, the concentrations of all other species must also change in order to maintain the same value for the equilibrium constant. This concept is stated by
Le Châtelier's Principle:

> *if a stress is applied to a system at chemical equilibrium, the equilibrium will shift is such a manner as to counteract the effects of that stress.*

A system (or reaction) may be "stressed" by a change in concentration of one or more species, a change in temperature, or a change in pressure. In the following series of experiments, you will observe and interpret the effects of changes in concentration and temperature upon systems at equilibrium.

Experimental Procedure

1. **Iron(III) ion with thiocyanate ion**
   a. Combine 2 drops of 0.2 M $Fe(NO_3)_3$ and 4 drops of 0.1 M KSCN in one well of the culture plate. (The reagents are preloaded in Beral pipets (plastic droppers).) Add 40 drops of deionized water and mix thoroughly. Note the colors of each reagent and of the product.
   b. The two reagents react to form a complex ion according to the reaction below:

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

Note: potassium and nitrate ions are colorless spectator ions.

c. Divide the solution among three separate wells of the plate. Be sure to note the color, as this will be used as your reference point for the original system, which contains all three species. (You may want to slide a piece of white paper under the culture plate to make the colors more distinguishable.)

d. To the first portion, add 5-8 drops of $Fe(NO_3)_3$, mixing after each drop, and observe all changes carefully. You may add more than 8 drops if necessary to observe a change).

e. To the second portion, add 5-8 drops of KSCN, mix, and observe.

f. To the third well, add 5-8 drops of 6 M NaOH, mix, and observe. The mixture may need to stand for a few minutes. Note: the compound iron(III) hydroxide, $Fe(OH)_3$ is quite insoluble in water.

2. Ammonium nitrate with water

a. Fill a large, dry test tube to a depth of about 2-3 cm with solid ammonium nitrate.

b. Add a volume of water equal to about one-half the volume of the solid and shake to mix. <u>Do not dissolve</u> all of the solid! Observe any temperature change (qualitatively). This system is described by the following reaction:

$$NH_4NO_3(s) + H_2O(l) \rightleftharpoons NH_4NO_3(aq) \qquad (4)$$

c. Heat may be regarded as a reactant or product in chemical reactions.

d. Heat is absorbed (reactant) in endothermic reactions; heat is evolved (product) in exothermic reactions.

e. Predict the effect on the system of raising the temperature. Test your prediction experimentally.

f. Is this reaction endothermic or exothermic?

3. Cobalt(II) ion, water, and chloride ion

In aqueous solution, $Co(NO_3)_2$ dissociates to form the pink octahedral complex ion $[Co(H_2O)_6]^{2+}$ (hexaaquacobalt(II) ion). However, in the presence of excess chloride ion, it is possible to form the tetrahedral complex ion $[CoCl_4]^{2-}$ (tetrachlorocobaltate(II) ion). The chemical equation for the reaction is shown below:

$$[Co(H_2O)_6]^{2+} + 4\ Cl^- \rightleftharpoons [CoCl_4]^{2-} + 6\ H_2O \qquad (5)$$

Concentration effects

a. Place 5 drops of 0.4 M $Co(NO_3)_2$ in an empty well of the culture plate. This solution exhibits the color of the hexaaquacobalt(II) ion (nitrate ion is colorless).

b. Half-fill another well with 12 M HCl (**Caution!** This solution is highly corrosive. If you spill it on yourself, rinse the area with copious amounts of water and inform your instructor.)

c. Use a Pasteur pipet to add 12 drops of the HCl to the $Co(NO_3)_2$ solution, mixing thoroughly and observing **after each drop**.

d. Now add 20 drops of water to the same well in small increments, mixing and observing as before.
e. Explain all observations.
f. Now add enough anhydrous calcium chloride to the same well to form a layer covering the bottom of the well. **Do not mix.**
g. The calcium chloride is a **dessicant** (look the word up) and should not completely dissolve.
h. Let the mixture stand for several minutes before recording your observations.
i. Explain what has happened (the chloride ion in the calcium chloride is not of significance).

Temperature effects

a. Place 5 mL of 0.4 M $CoCl_2$ into a small flask and add 3 mL of concentrated hydrochloric acid, mixing thoroughly.
b. The solution should be magenta. If it is not, add a few drops of water or HCl until a red-violet color is attained.
c. Note that hydrogen ion and chloride ion are colorless. The solution now contains appreciable amounts of both major cobalt complex ions in equation 5.
d. Divide the solution equally among three test tubes.
e. Maintain one test tube at room temperature.
f. Place the second test tube in a boiling water bath and place the third in an ice bath. Observe.
g. Allow the test tubes to return to room temperature, then switch the hot and cold test tubes . Are the changes reversible?