Equilibrium and Le Chatelier's Principle

Le Chatelier's Principle

- The French chemist Henri Le Chatelier (1850-1936) studied how the equilibrium position *shifts* as a result of changing conditions

- Le Chatelier's principle: If stress is applied to a system in equilibrium, the system changes in a way that relieves the stress

Le Chatelier's Principle

- What items are considered to be *stress* on the equilibrium?
  1) Concentration
  2) Temperature
  3) Pressure

Each of these will now be discussed in detail

What Le Chatelier found when studying equilibrium

Let's look at Cobalt Chloride

If we add water we get a cobalt complex that is a cation

$$CoCl_2\,(s) + 6H_2O \rightarrow \underset{\text{Pink}}{Co(H_2O)_6^{2+}} + 2Cl^-$$

at equilibrium

$$\underset{\text{Pink}}{Co(H_2O)_6^{2+}} + 4Cl^- \rightleftharpoons \underset{\text{Blue}}{CoCl_4^{2-}} + 6H_2O$$

at equilibrium

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

Pink — Blue

If we add HCL we get cobalt complex that is an anion

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

Blue

Let's see what the equilibrium constant looks like:

$$aA + bB \rightarrow cC + dD$$

$$K_{eq} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

$$Co(H_2O)_6^{2+} + 4Cl^- \rightarrow CoCl_4^{2-} + 6H_2O$$

Pink *Blue*

$$\mathbf{K_{eq} = \frac{[CoCl_4^{2-}]\ [H_2O]^6}{[Co(H_20)_6^{2+}]\ [Cl^-]^4}}$$

Effect of concentration change

$$K_{eq} = \frac{[CoCl_4^{2-}]}{[Co(H_20)_6^{2+}] [Cl^-]^4}$$

More of this forms

increases

Add hydrochloric acid (H^+) and (Cl^-)

From slide 5

$K_{eq} > 1$, products favored at equilibrium

$K_{eq} < 1$, reactants favored at equilibrium

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

Pink — Blue

Drive it back!

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

Blue

Increases

- Adding more water will drive the reaction in reverse

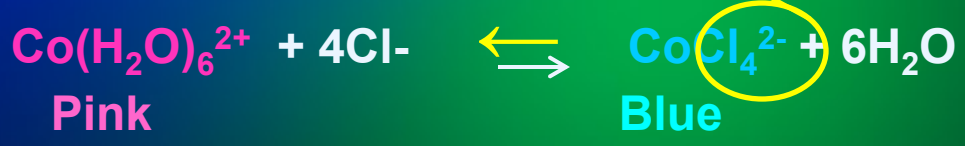
$Co(H_2O)_6^{2+} + 4Cl^- \leftarrow CoCl_4^{2-} + 6H_2O$

Pink

Now let's take Chlorine away and see what happens

Add Silver Nitrate

Ag^+ and Cl^- are strongly attracted to form an ion compound of AgCl. This removes Cl^- as $AgCl_{(s)}$



$$K_{eq} = \frac{[\text{now less than before}]}{[\text{now greater}]\ x\ [Cl^-]^4}$$

Pink

Effect of Concentration

- Concentration
  - – adding more reactant produces more product
  - – removing the product, as it forms, will produce more reactant

Effect of Temperature on Equilibrium

$Co(H_2O)_6^{2+} + 4Cl^- + \textbf{heat} \rightleftharpoons CoCl_4^{2-} + 6H_2O$

Pink *Blue*

To reverse the process, remove the heat

$Co(H_2O)_6^{2+} + 4Cl \rightleftharpoons CoCl_4^{2-} + 6H_2O$ **- heat**

Pink *Blue*

Le Chatelier's Principle and Temperature

- Temperature – increasing the temperature causes the equilibrium position to shift in the direction that absorbs heat

- If heat is one of the products (just like a chemical), it is part of the equilibrium

- so cooling an exothermic reaction will produce more product, and heating it would shift the reaction to the reactant side of the equilibrium

Le Chatelier's Principle and Pressure

- Pressure – changes in pressure will only effect gaseous equilibria
  - Increasing the pressure will usually favor the direction that has fewer molecules

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

- For every two molecules of ammonia made, four molecules of reactant are used up – the equilibrium *shifts to the right* with an increase in pressure