1.10.5.1: The Effect of a Concentration Change on Equilibrium

Consider the following system under equilibrium:

$$\underbrace{\mathrm{Fe^{3+}(aq)}}_{\text{colorless}} + \underbrace{\mathrm{SCN^{-}(aq)}}_{\text{colorless}} \rightleftharpoons \underbrace{\mathrm{FeSCN^{2+}(aq)}}_{\text{red}}$$

If more Fe^{3+} is added to the reaction, what will happen?

According to Le Chatelier's Principle, the system will react to minimize the stress. Since Fe^{3+} is on the reactant side of this reaction, the rate of the forward reaction will increase in order to "use up" the additional reactant. This will cause the equilibrium to **shift to the right**, producing more $FeSCN^{2+}$. For this particular reaction, we will be able to see that this has happened, as the solution will become a darker red color.

There are a few different ways to state what happens here when more Fe^{3+} is added, all of which have the same meaning:

- equilibrium shifts to the right
- equilibrium shifts to the product side
- the forward reaction is favored

What changes does this cause in the concentrations of the reaction participants?

Changes within reaction participants

$\mathrm{Fe^{3+}}$	Since this is what was added to cause the stress, the concentration of $\mathrm{Fe^{3+}}$ will increase. (A shorthand way to indicate this: $[\mathrm{Fe}]^{3+} \uparrow$ (Reminder: the square brackets represent "concentration")
$\mathrm{SCN^{-}(aq)}$	Equilibrium will shift to the right, which will use up the reactants. The concentration of $\mathrm{SCN^{-}(aq)}$ will decrease $[\mathrm{SCN}]^{-} \downarrow$ as the rate of the forward reaction increases.
$\mathrm{FeSCN^{2+}}$	When the forward reaction rate increases, more products are produced, and the concentration of $\mathrm{FeSCN^{2+}}$ will increase. $[\mathrm{FeSCN}]^{2+} \uparrow$

How about the value of K_{eq}? Notice that the concentration of some reaction participants have increased, while others have decreased. Once equilibrium has re-established itself, the value of K_{eq} will be unchanged.

> *The value of K_{eq} does not change when changes in concentration cause a shift in equilibrium.*

What if more $FeSCN^{2+}$ is added?

Again, equilibrium will shift to use up the added substance. In this case, equilibrium will shift to favor the *reverse* reaction, since the reverse reaction will use up the additional $FeSCN^{2+}$.

- equilibrium shifts to the left
- equilibrium shifts to the reactant side
- the reverse reaction is favored

How do the concentrations of reaction participants change?

Change of concentrations of reaction participants when adding substance

$\mathrm{Fe^{3+}}$	$[\mathrm{Fe}]^{3+} \uparrow$ as the reverse reaction is favored
$\mathrm{SCN^{-}(aq)}$	$[\mathrm{SCN}]^{-} \uparrow$ as the reverse reaction is favored
$\mathrm{FeSCN^{2+}}$	$[\mathrm{FeSCN}]^{2+} \uparrow$ because this is the substance that was added

Concentration can also be changed by **removing** a substance from the reaction. This is often accomplished by adding another substance that reacts (in a side reaction) with something already in the reaction.

Let's remove SCN^- from the system (perhaps by adding some Pb^{2+} ions—the lead(II) ions will form a precipitate with SCN^-, removing them from the solution). What will happen now? Equilibrium will shift to **replace** SCN^-—the reverse reaction will be favored because that is the direction that produces more SCN^-.

- equilibrium shifts to the left
- equilibrium shifts to the reactant side
- the reverse reaction is favored

How do the concentrations of reaction participants change?

Change of concentrations of reaction participants when removing a substance

Fe^{3+}	$[Fe]^{3+}$ ↑ as the reverse reaction is favored
SCN^-	$[SCN]^-$ ↑ as the reverse reaction is favored (but also ↓ because it was removed)
$FeSCN^{2+}$	$[FeSCN]^{2+}$ ↑ because this is the substance that was added