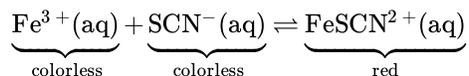


## 1.10.5.1: The Effect of a Concentration Change on Equilibrium

Consider the following system under equilibrium:



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

According to Le Chatelier's Principle, the system will react to minimize the stress. Since  $\text{Fe}^{3+}$  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  $\text{FeSCN}^{2+}$ . 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  $\text{Fe}^{3+}$  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

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

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

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

What if more  $\text{FeSCN}^{2+}$  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  $\text{FeSCN}^{2+}$ .

- 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

$\text{Fe}^{3+}$	$[\text{Fe}]^{3+} \uparrow$ as the reverse reaction is favored
$\text{SCN}^{-}(\text{aq})$	$[\text{SCN}]^{-} \uparrow$ as the reverse reaction is favored
$\text{FeSCN}^{2+}$	$[\text{FeSCN}]^{2+} \uparrow$ 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  $\text{SCN}^-$  from the system (perhaps by adding some  $\text{Pb}^{2+}$  ions—the lead(II) ions will form a precipitate with  $\text{SCN}^-$ , removing them from the solution). What will happen now? Equilibrium will shift to **replace**  $\text{SCN}^-$ —the reverse reaction will be favored because that is the direction that produces more  $\text{SCN}^-$ .

- 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

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

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