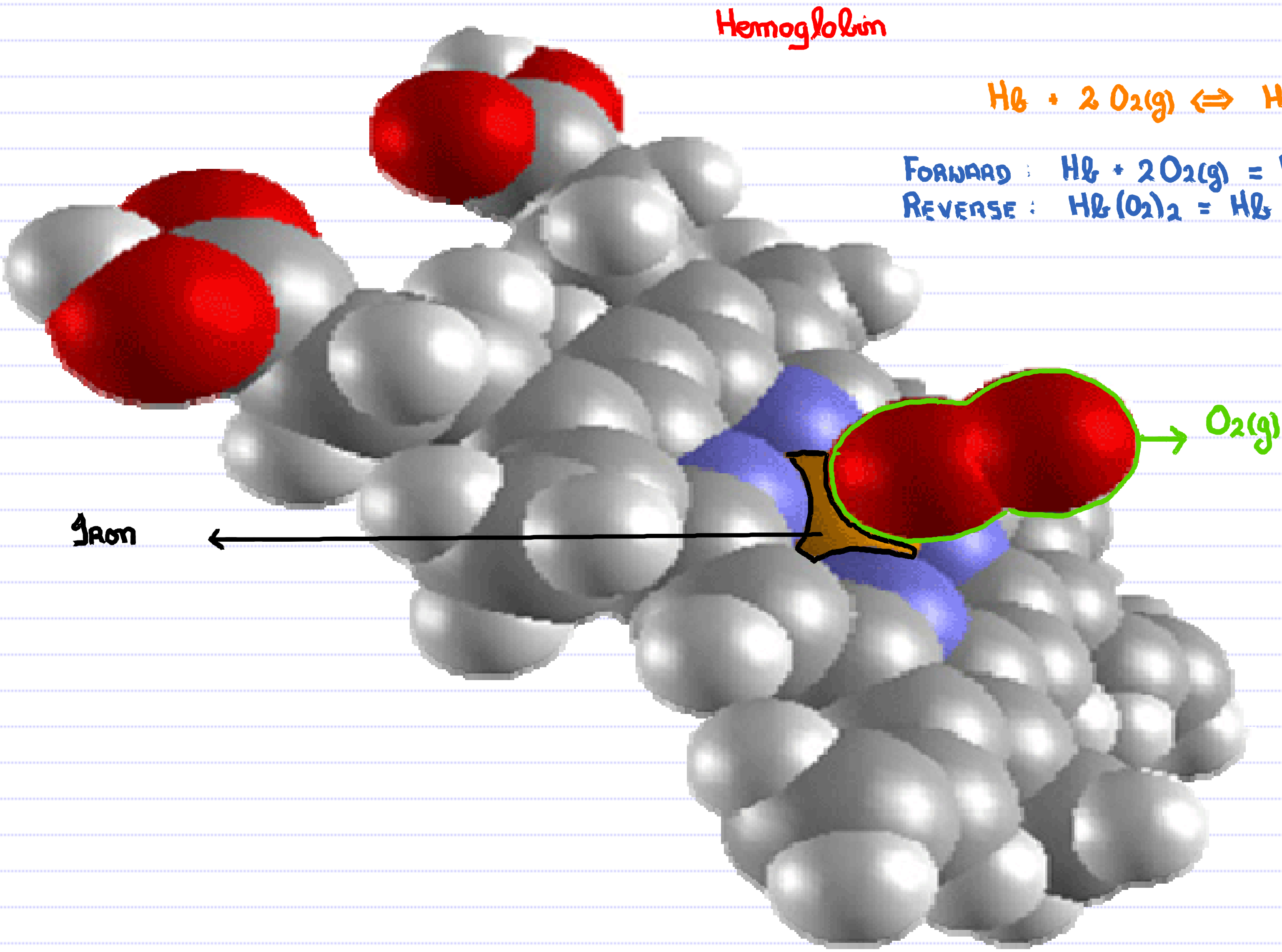
15.1 **The Nature of the Equilibrium State**

The Equilibrium State



$Hb + 2\,O_2(g) \Leftrightarrow Hb(O_2)_2$

FORWARD : $Hb + 2O_2(g) = Hb(O_2)_2$

REVERSE : $Hb(O_2)_2 = Hb + 2O_2(g)$

15.1 The Nature of the Equilibrium State

The Equilibrium State

$$Fe^{3+} + SCN^- \Leftrightarrow FeSCN^{2+}$$

At Equilibrium : Initial Rate of forward Reaction = Initial Rate of the Reverse Reaction.

$$k_f [Fe^{3+}][SCN^-] = k_R [FeSCN^{2+}]$$

$$\frac{k_f}{k_R} = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$$

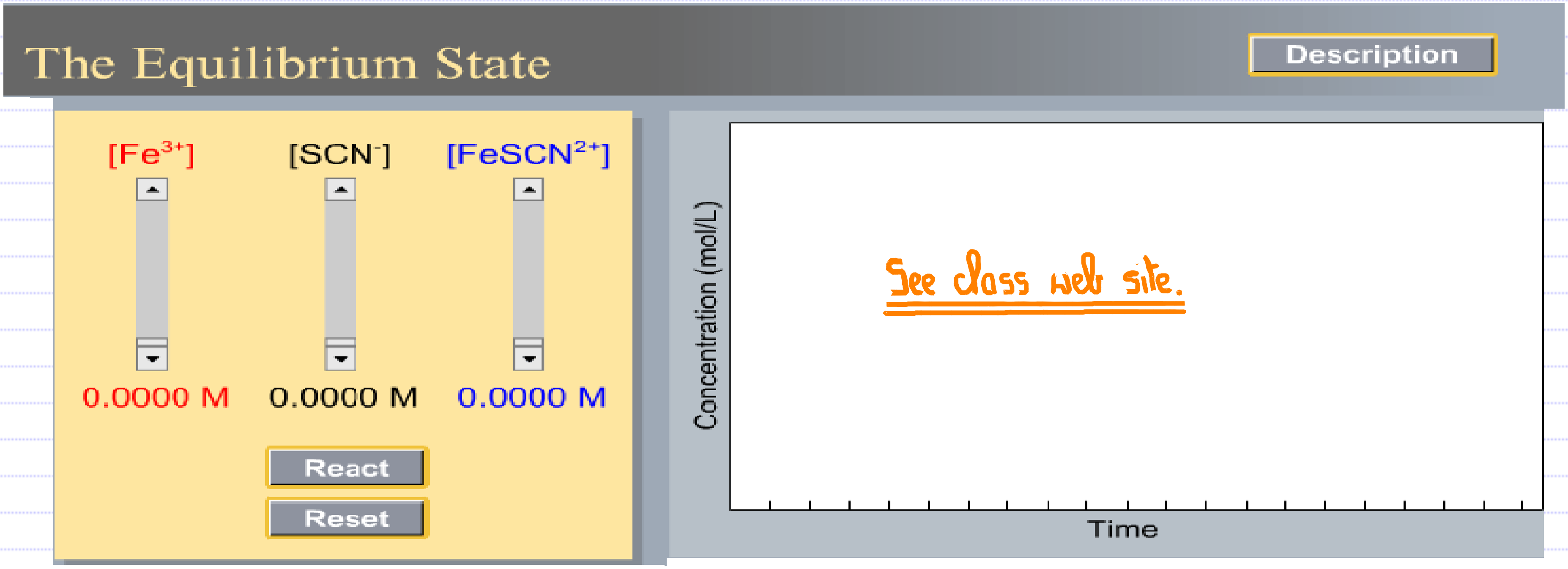
→ Constant.

$$\frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]} = \text{Constant} = K$$

? Is this true ... don't take my word for it ! ... experiment

15.2 The Equilibrium Constant, K

The Equilibrium State

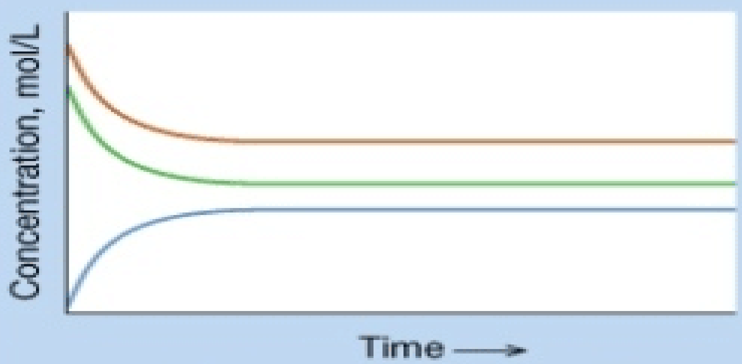


$Fe^{3+} + SCN^{-} \rightleftharpoons FeSCN^{2+}$

	Initial Concentrations	Final Concentrations
$[Fe^{3+}]$ =	0.0000 M	0.0000 M
$[SCN^{-}]$ =	0.0000 M	0.0000 M
$[FeSCN^{2+}]$ =	0.0000 M	0.0000 M

15.2 The Equilibrium Constant, K

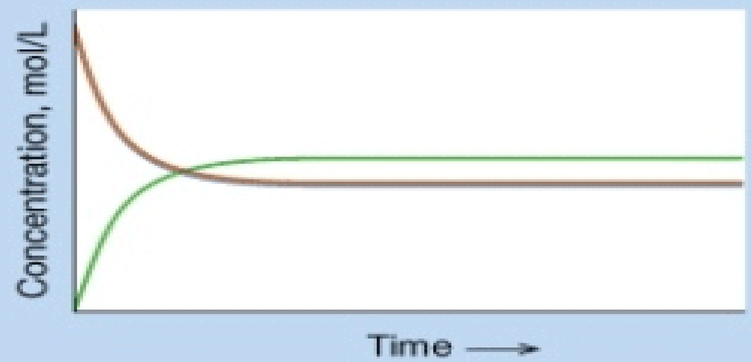
Equilibrium Constants



Initial concentrations	Final concentrations
$[Fe^{3+}] = 0.00900$ M	5.63×10^{-3} M
$[SCN^{-}] = 0.00760$ M	4.23×10^{-3} M
$[FeSCN^{2+}] = 0.00000$ M	3.37×10^{-3} M

Fe^{3+} and SCN^{-}

$$K = \frac{[0.00337]}{[0.00563]\,[0.00423]} = 142$$



Initial concentrations	Final concentrations
$[Fe^{3+}] = 0.00000$ M	5.18×10^{-3} M
$[SCN^{-}] = 0.00000$ M	5.18×10^{-3} M
$[FeSCN^{2+}] = 0.00900$ M	3.82×10^{-3} M

Fe^{3+} and SCN^{2+}

$$K = \frac{[0.00382]}{[0.00518]\,[0.00518]} = 142$$

In general: $aA + bB \Leftrightarrow cC + dD$

$$K = \frac{[\text{PRODUCTS}]}{[\text{REACTANTS}]} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

15.2 The Equilibrium Constant, K

Equilibrium Constants – Meaning of the Magnitude of K

K : $K >> 1$; $K << 1$; $K \approx 1$

$K >> 1$

$I_2(g) + Cl_2(g) \Leftrightarrow 2\,ICl(g)$ $K = 2.1 \times 10^5$ @ 25°C

Product-favored reaction.

@ Equilibrium :- very little $I_2(g)$ and $Cl_2(g)$ remaining.

$K << 1$

$CH_3CO_2H(aq) + H_2O(l) \Leftrightarrow CH_3CO_2^- + H_3O^+$ $K = 1.8 \times 10^{-5}$ @ 25°C

Reactant-favored reaction.

@ Equilibrium :- very little $CH_3CO_2^-$ and H_3O^+ produced.

$K \approx 1$

$2\,NO_2(g) \Leftrightarrow N_2O_4(g)$ $K = 1.4$ @ 25°C

@ Equilibrium :- significant amounts of $NO_2(g)$ and $N_2O_4(g)$ present.

15.2 The Equilibrium Constant, K

Equilibrium Constants – Meaning of the Magnitude of K

The equilibrium constant, Kc, for the following reaction is 1.29×10^{-6} at 600 K.

$$COCl_2(g) \Leftrightarrow CO(g) + Cl_2(g)$$

Assuming that you start with only $COCl_2$, describe the relative abundance of each species present at equilibrium.

$[COCl_2(g)]$	a. Higher ✓	b. Lower	c. Can't tell
$[CO(g)]$	a. Higher	b. Lower ✓	c. Can't tell

$K = 1.29 \times 10^{-6}$ @ 600K

→ Reactant-favored reaction.

15.2 The Equilibrium Constant, K

Writing Equilibrium Constant Expressions

$$aA + bB \rightleftharpoons cC + dD$$

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

However: a) Pure solids do not appear in the expression.

b) Pure liquids and solvents do not appear in the expression.

a) $C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$ $K = \frac{[H_2][CO]}{[H_2O]}$

b) $CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons CH_3CO_2^- + H_3O^+$ $K = \frac{[CH_3CO_2^-][H_3O^+]}{[CH_3CO_2H]}$