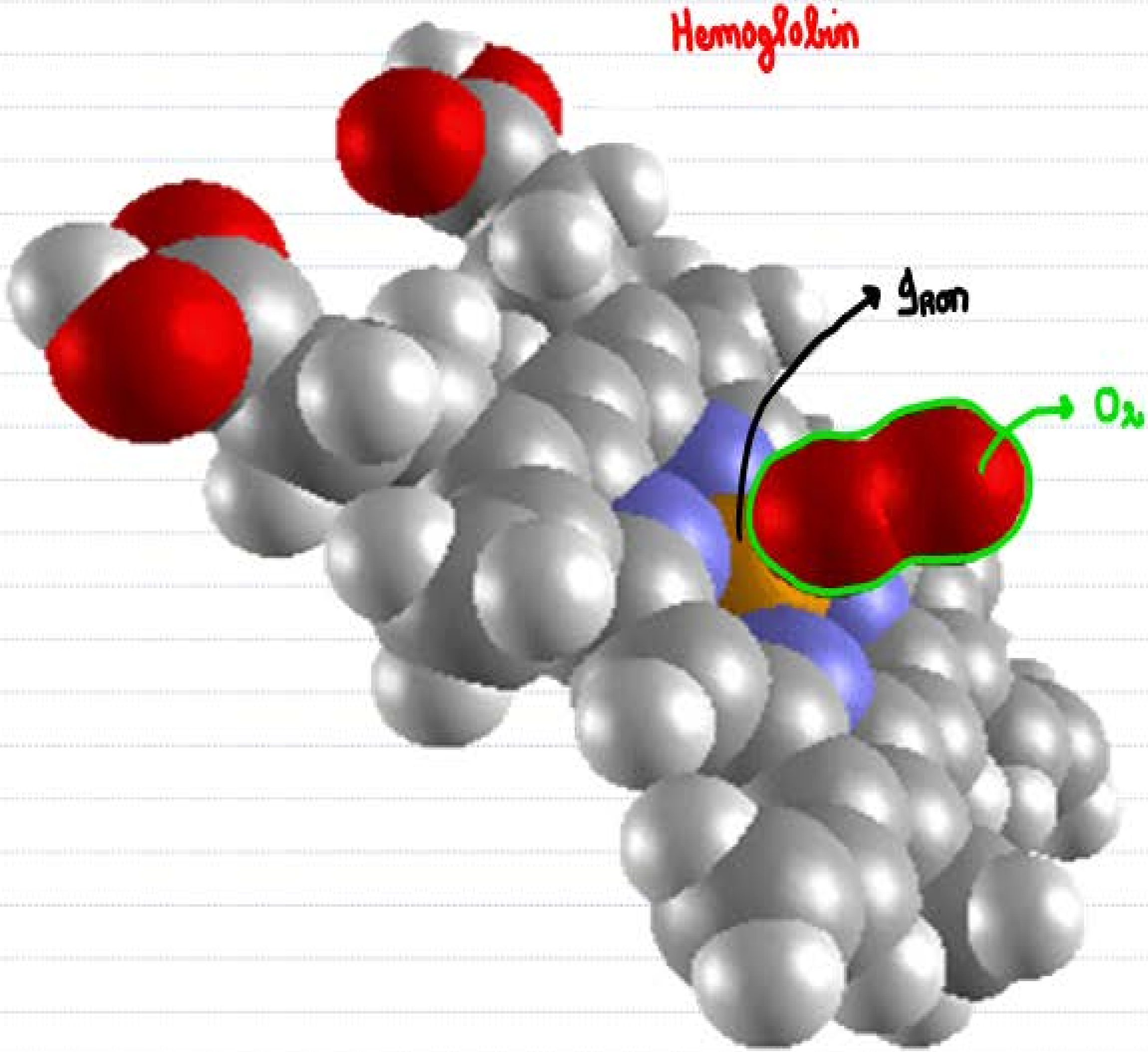
Announcements – Lecture X – Tuesday, Feb 27th

1. iClicker:  Pick any letter a-e

15.1 The Nature of the Equilibrium State

The Equilibrium State



$HbO_2 \Leftrightarrow Hb + O_2$

forward : $HbO_2 = Hb + O_2$

reverse : $Hb + O_2 = HbO_2$

15.1 The Nature of the Equilibrium State

The Equilibrium State

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

At equilibrium : Rate of the forward reaction = Rate of the reverse reaction

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

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

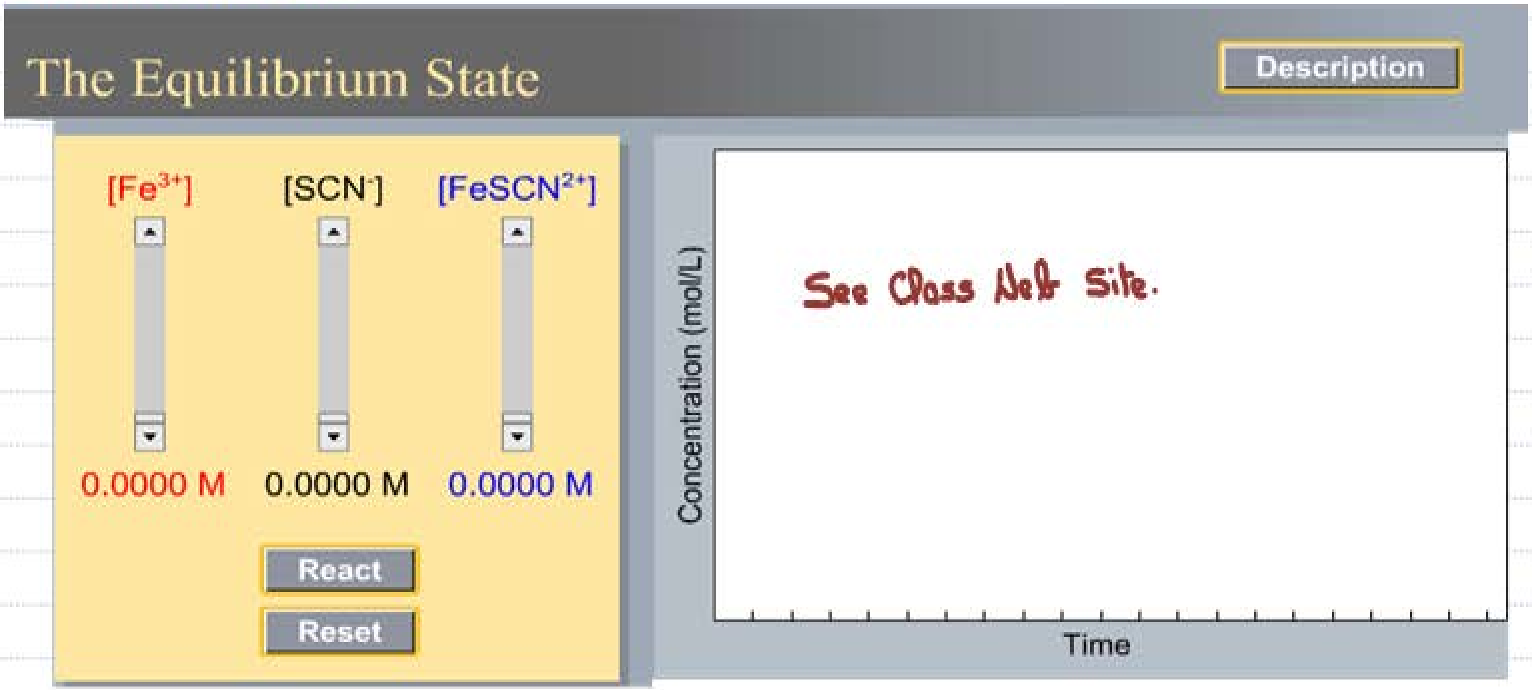
↳ constant

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

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

15.2 The Equilibrium Constant, K

The Equilibrium State

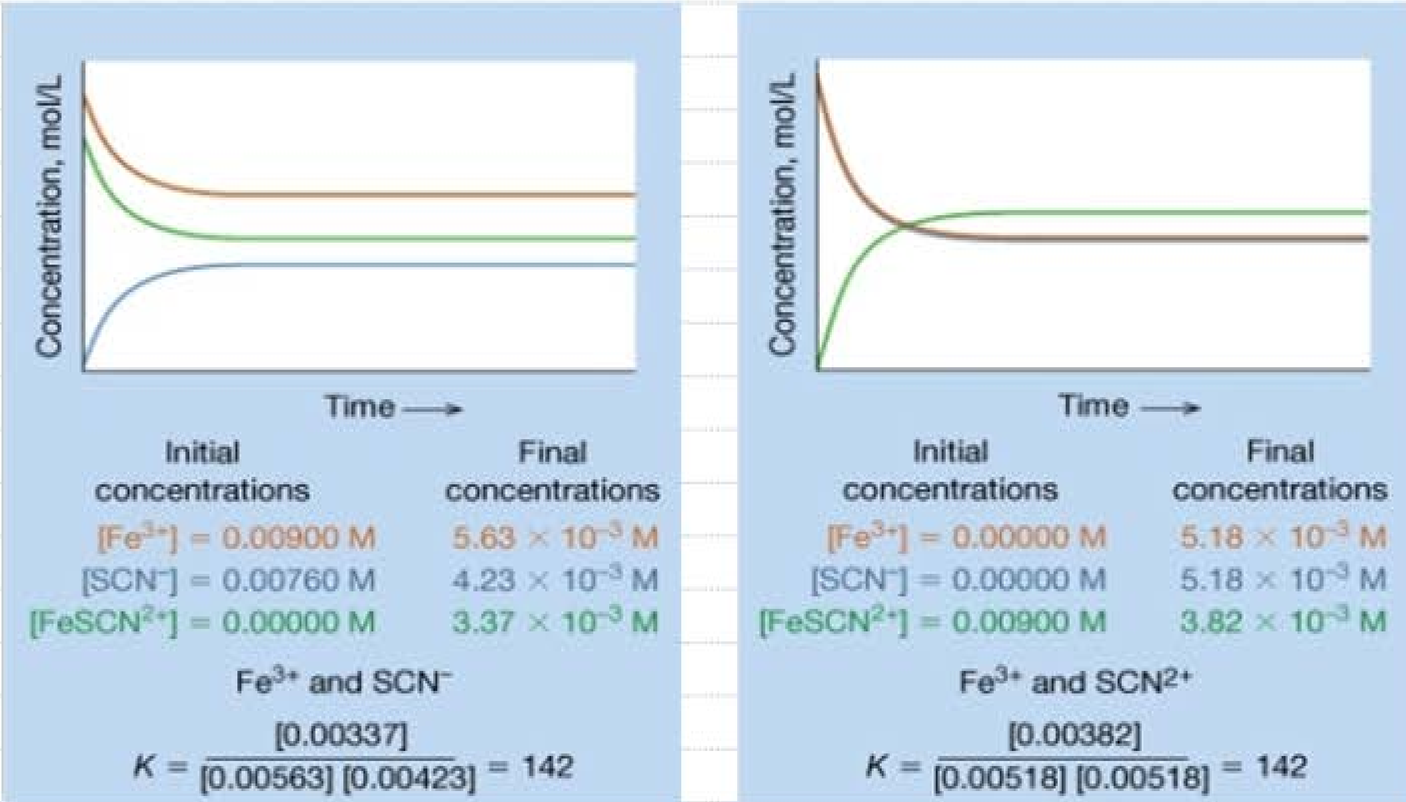


$Fe^{3+} + SCN^{-} \Leftrightarrow FeSCN^{2+}$

	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



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: $K \gg 1$, $K \ll 1$, $K \approx 1$

$K \gg 1$: $I_2(g) + Cl_2(g) \Leftrightarrow 2\,ICl(g)$ $K = 2.1 \times 10^5$ @ 25°C

product-favored reaction.

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

$K \ll 1$: $CH_3CO_2H(aq) + H_2O(l) \Leftrightarrow CH_3CO_2^- + H_3O^+$ $K = 1.8 \times 10^{-5}$ @ 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

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.

$[COCl_2(g)]$ 1. Higher ✓ 2. Lower 3. Can't tell

$[CO(g)]$ 1. Higher 2. Lower ✓ 3. Can't tell

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

Reactant-favored reaction.

15.2 The Equilibrium Constant, K

Writing Equilibrium Constant Expressions

$$aA + bB \Leftrightarrow cC + dD$$

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

However : i) Pure solids do not appear in the expression.
ii) Pure liquids and solvents do not appear in the expression.

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

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

15.2 The Equilibrium Constant, K

The Relationship between K_p and K_c

$$2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$$

$$K_c = \frac{[NO]^2[Br_2]}{[NOBr]^2} \qquad K_p = \frac{P_{NO}^2 P_{Br_2}}{P_{NOBr}^2}$$

How are K_c and K_p related?

$$PV = nRT$$

$$P = (n/V)RT$$

$\hookrightarrow$ mol.ℓ$^{-1}$ = [ ] RT

$$K_p = \frac{P_{NO}^2 P_{Br_2}}{P_{NOBr}^2}$$

$$= \frac{[NO]^2(RT)^2[Br_2](RT)}{[NOBr]^2(RT)^2}$$

$$= \frac{[NO]^2[Br_2]}{[NOBr]^2} \times \frac{(RT)^3}{(RT)^2}$$

$$K_p = K_c(RT)^3(RT)^{-2}$$

$$= K_c(RT)^{3-2}$$

3 - 2 = mol gas products - mol gas reactants

$2\,NO(g) + Br_2(g)$ — $2\,NOBr(g)$

3 - 2

$$K_p = K_c(RT)^{\Delta n}$$

Δn = mol gas products - mol gas reactants