Equilibrium Calculations

A problem

- $2\ SO_{2\ (g)} + O_{2\ (g)} \rightleftharpoons 2\ SO_{3\ (g)}$
- K = 4.34, for the above reaction. Calculate the concentration of SO_3 if the SO_2 = .28 M and O_2 = .43 M at equilibrium.

Answer

- $K = \dfrac{[SO_3]^2}{[SO_2]^2\,[O_2]}$

- $4.34 = \dfrac{[SO_3]^2}{[.28]^2[.43]}$

- $[SO_3]$ = .38 M

Reaction Quotient , Q

- The reaction quotient, Q, gives us a value of **current concentrations** of all chemicals before they complete reacting towards equilibrium.
  - It is calculated the same way as K except you use current concentrations.
  - Comparing the Q value to the K value allows you to determine which way the reaction will shift.

K and Q values

K and Q are calculated by products over reactants.

- Larger values mean higher products, smaller mean higher reactants.
- If **K is larger than Q**, it means that Q has a higher concentration of reactants so the reaction will shift to the **right**.
- If **K is smaller than Q**, it means that Q has a higher concentration of products so the reaction will shift to the **left**.
- If **K equals Q** you are at equilibrium, **no shift**

Easy way to remember

- Greater than, less than signs are an arrow!!!
- Put K first, then Q and determine which is larger.
- $K > Q$, shift right
- $K < Q$, shift left
- $K = Q$, No shift

Q problem

- The K is .060 for the reaction:
- $N_2 + 3\ H_2 \rightleftharpoons 2\ NH_3$
- Calculate Q and determine which way the reaction will shift

	$[N_2]$ (M)	$[H_2]$ (M)	$[NH_3]$(M)
Case 1	1.0×10^{-5}	2.0×10^{-3}	1.0×10^{-3}
Case 2	1.5×10^{-5}	3.54×10^{-1}	2.0×10^{-4}
Case 3	5.0	1.0×10^{-2}	1.0×10^{-4}

Answer

- $Q = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$
- Case 1, $Q = 1.25 \times 10^7$
- K < Q, reaction shifts left
- Case 2, Q = .060
- K = Q, equilibrium
- Case 3, Q = .002
- K> Q, reaction shifts right

Solution Equilibrium

- All dissociations we have done are equilibriums.
- Before we simply stated something was soluble or insoluble.
- Actually everything dissolves to some extent, and some dissolved substance fall out of solution.
- Higher concentrations force more solute to fall out of solution.
- So there is a maximum concentration of solute a solution can hold (saturation)

Solution equilibrium

- For Example:
- Lead (II) Bromide
- $PbBr_{2\,(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2\ Br^{-}{}_{(aq)}$
- What would the equilibrium expression look like?
- $K_{sp} = [Pb^{2+}][Br^{-}]^2$
- Equilibrium constants for dissociations are called **solubility products**, and are denoted by $\mathbf{K_{sp}}$.

Cont.

The K_{sp} value for $PbBr_2$ is 5.0×10^{-6}.
$PbBr_{2\,(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2\ Br^{-}{}_{(aq)}$

- Say the maximum moles of $PbBr_2$ that can dissolve is x moles per 1L.
- Using the balanced equation, I will have x mole/L of Pb^{2+} and 2x mole/L of Br^-
- So
- $5.0 \times 10^{-6} = x\,(2x)^2$
- $5.0 \times 10^{-6} = 4x^3$
- x=.011 M This is the solubility, or maximum amount of solute, $PbBr_2$, that will dissolve.
- $[Pb^{2+}]$ = .011 M 2x = $[Br^-]$= .022 M

This means

- A solution of $PbBr_2$ would be saturated with $[PbBr_2]$ = .011 M
- After I reach that level, no more will dissolve.
- This is a very low concentration, so we say the is compound is insoluble.
- In the solution, the ion concentrations would be
- $[Pb^{2+}]$ = .011 M
- $[Br^-]$= .022 M

Problems

- Calculate the solubility and saturation concentrations of ions in aluminum hydroxide K_{sp} = 5.0 x10 $^{-33}$, and Barium sulfate K_{sp} = 1.4 x10 $^{-14}$

Answer

- $Al(OH)_3 \rightleftharpoons Al^{3+}{}_{(aq)} + 3\ OH^-{}_{(aq)}$
- $K_{sp} = [Al^{3+}][OH^-]^3$
- $5.0\times10^{-33} = x\,(3x)^3$
- $x = 3.7\times10^{-9}$ M (solubility)

- $[Al^{3+}] = 3.7\times10^{-9}$ M
- $3x = [OH^-] = 1.1\times10^{-8}$ M

Answer

- $BaSO_4 \rightleftharpoons Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$
- $K_{sp} = [Ba^{2+}][SO_4^{2-}]$
- $1.4\times10^{-14} = x\,(x)$
- $x = 1.2\times10^{-7}$ M (solubility)

- $[Ba^{2+}] = 1.2\times10^{-7}$ M
- $[SO_4^{2-}] = 1.2\times10^{-7}$ M