

Equilibrium Calculations

A problem

- $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$
- $K = 4.34$, for the above reaction.
Calculate the concentration of SO_3 if the $\text{SO}_2 = .28 \text{ M}$ and $\text{O}_2 = .43 \text{ M}$ at equilibrium.

Answer

- $K = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$
- $4.34 = \frac{[\text{SO}_3]^2}{[.28]^2[.43]}$
- $[\text{SO}_3] = .38 \text{ M}$

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

- $Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$
- 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

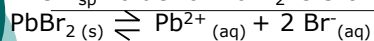
- All dissociations we have done are equilibria.
- Before we simply stated something was soluble or insoluble.
- Actually everything dissolves to some extent, and some dissolved substance fall out of solution.
- Higher concentrations force more solute to fall out of solution.
- So there is a maximum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

Solution equilibrium

- For Example:
- Lead (II) Bromide
- $\text{PbBr}_2 (s) \rightleftharpoons \text{Pb}^{2+} (aq) + 2 \text{Br}^- (aq)$
- What would the equilibrium expression look like?
- $K_{sp} = [\text{Pb}^{2+}][\text{Br}^-]^2$
- Equilibrium constants for dissociations are called **solubility products**, and are denoted by **K_{sp}** .

Cont.

The K_{sp} value for PbBr_2 is 5.0×10^{-6} .



- 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
- $5.0 \times 10^{-6} = x(2x)^2$
- $5.0 \times 10^{-6} = 4x^3$
- $x = .011 \text{ M}$ This is the solubility, or maximum amount of solute, PbBr_2 , that will dissolve.
- $[\text{Pb}^{2+}] = .011 \text{ M}$ $2x = [\text{Br}^-] = .022 \text{ M}$

This means

- A solution of PbBr_2 would be saturated with $[\text{PbBr}_2] = .011 \text{ M}$
- After I reach that level, no more will dissolve.
- This is a very low concentration, so we say the compound is insoluble.
- In the solution, the ion concentrations would be
- $[\text{Pb}^{2+}] = .011 \text{ M}$
- $[\text{Br}^-] = .022 \text{ M}$

Problems

- Calculate the solubility and saturation concentrations of ions in aluminum hydroxide $K_{sp} = 5.0 \times 10^{-33}$, and Barium sulfate $K_{sp} = 1.4 \times 10^{-14}$

Answer

- $\text{Al}(\text{OH})_3 \rightleftharpoons \text{Al}^{3+}_{(\text{aq})} + 3 \text{OH}^{-}_{(\text{aq})}$
- $K_{sp} = [\text{Al}^{3+}][\text{OH}^{-}]^3$
- $5.0 \times 10^{-33} = x(3x)^3$
- $x = 3.7 \times 10^{-9} \text{ M (solubility)}$

- $[\text{Al}^{3+}] = 3.7 \times 10^{-9} \text{ M}$
- $3x = [\text{OH}^{-}] = 1.1 \times 10^{-8} \text{ M}$

Answer

- $\text{BaSO}_4 \rightleftharpoons \text{Ba}^{2+}_{(\text{aq})} + \text{SO}_4^{2-}_{(\text{aq})}$
- $K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$
- $1.4 \times 10^{-14} = x(x)$
- $x = 1.2 \times 10^{-7} \text{ M (solubility)}$

- $[\text{Ba}^{2+}] = 1.2 \times 10^{-7} \text{ M}$
- $[\text{SO}_4^{2-}] = 1.2 \times 10^{-7} \text{ M}$
