Equilibrium: Reaction Quotients

Read from **Lesson 2d: Predicting the Direction of Reaction** in the **Chemistry Tutorial Section, Chapter 14** of **The Physics Classroom**.

Reaction Quotients

The **reaction quotient (Q)** is a number that shows the ratio of the concentrations (or partial pressures) of products to reactants at any point in a chemical reaction, not necessarily at equilibrium. It is calculated using the same expression as the equilibrium constant (K), but with the current concentrations or pressures.

By comparing Q to K, we can predict which direction the reaction will shift to reach equilibrium:

- If **Q < K**, the reaction will shift **forward** (toward products).
- If **Q > K**, the reaction will shift **backward** (toward reactants).
- If **Q = K**, the reaction is **at equilibrium**.

Example:
At a specific temperature, the reaction, $N_2\ (g) + 3\ H_2\ (g) \rightleftarrows 2\ NH_3\ (g)$ has an equilibrium constant: $K_c = 1.2$. Suppose that the following concentrations are known: $[N_2] = 0.50$ M, $[H_2] = 0.30$ M, and $[NH_3] = 0.10$ M. Is the system currently at equilibrium? If not, how does the system need to change to reach equilibrium?

First, find Q:

$$Q = \frac{[NH_3]^2}{[N_2] \bullet [H_2]^3} \qquad Q = \frac{[NH_3]^2}{[N_2] \bullet [H_2]^3} = \frac{(0.10\ M)^2}{(0.50\ M) \bullet (0.30\ M)^3} = 0.74$$

Reactants are favvored

- Since Q = 0.74 and K = 1.2, and Q < K, the reaction is not at equilibrium.
- Because Q is smaller than K, the reaction will shift to the right (toward products) to reach equilibrium.

Questions

1. When the reaction $2NOBr\ (g) \rightleftarrows 2\ NO\ (g) + Br_2\ (g)$ occurs at 400 K, the $K_p = 0.00314$. If the partial pressures of the system are known to be $P_{NOBr} = 1.00$ atm, $P_{NO} = 1.00$ atm, and $P_{Br2} = 2.40$ atm.

    a. Will the reaction shift left, or right? Show work to explain your answer.

    b. Which partial pressure will increase as equilibrium is established?

Kinetics and Equilibrium

2 Aaron Agin is studying this reaction, $2\ SO_2\ (g) + O_2\ (g) \rightleftarrows 2\ SO_3(g)$, in the chem lab. The K_c at the lab temperature is 4.0. Aaron mixes the following solutions: $[SO_2] = 0.30$ M, $[O_2] = 0.15$ M, and $[SO_3] = 0.80$ M.

a. Aaron believes the reaction is at equilibrium. Is he correct? Show work to prove your answer.

b. Is this system at equilibrium? If not, which direction will the system proceed to reach equilibrium?

3. Aaron Agin is back again—this time discussing the reverse reaction with the identical twins, Flo and Clo Wrene. They're examining the equilibrium reaction: $2\ SO_3(g) \rightleftarrows 2\ SO_2\ (g) + O_2\ (g)$ with an equilibrium constant $K_P = 45$ at the lab temperature.
A sealed flask is filled with a mixture of all three gases. The initial pressures are $P_{SO2} = 0.500$ atm, $P_{O2} = 0.400$ atm, and $P_{SO3} = 1.00$ atm. After mixing the gases, Flo predicts that more SO_2 and O_2 will form. Clo argues that more SO_3 will form. Aaron claims the system is already at equilibrium. Who is correct? Explain why and explain what error the incorrect students are likely making.

4. For the reaction: $N_2\ (g) + O_2\ (g) \rightleftarrows 2\ NO\ (g)$ and $K_c = 14.36$. If the concentrations are known to be $[N_2] = 2.50$ M, $[O_2] = 3.50$ M, and $[NO] = 5.25$ M.

a. Is this system at equilibrium? If not, which direction will the system proceed to reach equilibrium?

b. How will the change in direction (if any) affect each of the concentrations? If changed, describe how each is different.