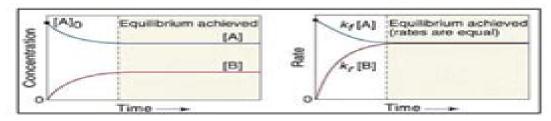
Chapter 14 Worksheet 1 (ws14.1)

Chemical Equilibrium, the Equilibrium Constant (Keq), and the Reaction Quotient (Q)

As a reaction proceeds in a **closed system**, the concentrations of reactants and products change until, at **equilibrium**, the rate of the forward reaction is equal to the rate of the reverse reaction. At that point there is **no net change** in the concentrations of reactants or products. For example, for the simple reaction: A = B:



Regardless of the initial concentrations, the concentrations of reactants and products present at equilibrium are related through the equilibrium constant (K_{eq}) as follows:

For the reaction: $aA + bB \Rightarrow cC + dD$

If all reactants and products are in solution then:

$$K_{eq} = K_e = \frac{[C]_{eq}^e[D]_{eq}^d}{[A]_{eq}^a[B]_{eq}^b}$$

(The subscript "eq" indicates "equilibrium" concentration. The subscript "c" stands for "concentration". The exponents are the coefficients from the balanced equation.)

If all reactants and products are ideal gases then either partial pressures or concentrations can be used in the equilibrium constant expression. In terms of partial pressures:

$$K_{eq} = K_p = \frac{(P_C)_{eq}^c (P_D)_{eq}^d}{(P_A)_{eq}^b (P_B)_{eq}^b}$$
 (The subscript "p" stands for "pressure".)

The numerical value of the two equilibrium constants above will be different. Remember the ideal gas law shows that the partial pressure of a gas is directly proportional to its molarity:

$$P = \frac{nRT}{V} = MRT \ \, (M \ \, \text{is molarity and} \, \, R = 0.0821 \ \, L\text{-atm/mol-K}).$$

Using this relationship, it is easy to show that (see pages 605-606 in your textbook):

$$K_p = K_c(RT)^{\Delta n}$$
 where $\Delta n =$ moles of gaseous products – moles of gaseous reactants

For reactions involving both gases and aqueous solutions, molar concentrations and partial pressures can appear in the equilibrium constant expression.

Chemical Equilibrium Review Answer Key

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