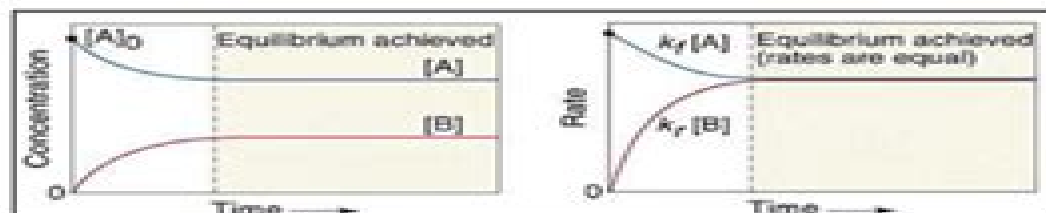


Chapter 14 Worksheet 1 (ws14.1)

Chemical Equilibrium, the Equilibrium Constant (K_{eq}), 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: $A \rightleftharpoons B$:



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: $aA + bB \rightleftharpoons cC + dD$

If all reactants and products are in solution then:

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

(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:

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

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:

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

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

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

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

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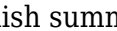
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