Section 7.2: Equilibrium Law and the Equilibrium Constant

Tutorial 1 Practice, page 431

1. (a) $2\ CO_2(g) \rightleftharpoons 2\ CO(g) + O_2(g)$

Products: $2\ CO(g)$; $O_2(g)$

Reactant: $2\ CO_2(g)$

Equilibrium law equation: $K = \dfrac{[CO(g)]^2[O_2(g)]}{[CO_2(g)]^2}$

(b) $2\ Cl_2(g) + 2\ H_2O(g) \rightleftharpoons 4\ HCl(g) + O_2(g)$

Products: $4\ HCl(g)$; $O_2(g)$

Reactants: $2\ Cl_2(g)$; $2\ H_2O(g)$

Equilibrium law equation: $K = \dfrac{[HCl(g)]^4[O_2(g)]}{[Cl_2(g)]^2[H_2O(g)]^2}$

(c) $2\ O_3(g) \rightleftharpoons 3\ O_2(g)$

Product: $3\ O_2(g)$

Reactant: $2\ O_3(g)$

Equilibrium law equation: $K = \dfrac{[O_2(g)]^3}{[O_3(g)]^2}$

(d) $4\ NH_3(g) + 3\ O_2(g) \rightleftharpoons 2\ N_2(g) + 6\ H_2O(g)$

Products: $2\ N_2(g)$; $6\ H_2O(g)$

Reactants: $4\ NH_3(g)$; $3\ O_2(g)$

Equilibrium law equation: $K = \dfrac{[N_2(g)]^2[H_2O(g)]^6}{[NH_3(g)]^4[O_2(g)]^3}$

2. Given: $CO(g) + 2\ H_2(g) \rightleftharpoons CH_3OH(g)$; $[CO(g)]_{equilibrium} = 0.079$ mol/L; $[H_2(g)]_{equilibrium} = 0.158$ mol/L; $[CH_3OH(g)]_{equilibrium} = 0.021$ mol/L

Required: K

Solution: Write the equilibrium law equation using the balanced chemical equation. Then, substitute the equilibrium concentrations into the equilibrium law equation and solve for K.

$$K = \frac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$$

$$= \frac{(0.021)}{(0.079)(0.158)^2}$$

$$K = 11$$

Statement: The equilibrium constant, K, for the reaction is 11.

3. Given: $K = \dfrac{[NO(g)]^2[Br_2(g)]}{[NOBr(g)]^2}$

Required: the balanced chemical equation for the reaction

Analysis: Use the general equilibrium law equation to determine the balanced reaction equation.

Solution: Products: 2 $NO(g)$; $Br_2(g)$

Reactant: 2 $NOBr(g)$

Balanced chemical equation: $2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$

Statement: The balanced chemical equation is $2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$.

Tutorial 2 Practice, page 434

1. (a) $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$

Products: $NH_3(g)$; $HCl(g)$

Reactant: $NH_4Cl(s)$

Equilibrium law equation: $K = [NH_3(g)][HCl(g)]$

(b) Unbalanced chemical equation: $NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$

Balanced chemical equation: $2\ NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$

Products: $Na_2CO_3(s)$; $CO_2(g)$; $H_2O(g)$

Reactant: 2 $NaHCO_3(s)$

Equilibrium law equation: $K = [CO_2(g)][H_2O(g)]$

(c) Chemical equation: $Zn(s) + Cu^{2+}(aq) + 2\ Cl^{-}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s) + 2\ Cl^{-}(aq)$

Net ionic equation: $Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$

Products: $Zn^{2+}(aq)$; $Cu(s)$

Reactants: $Zn(s)$; $Cu^{2+}(aq)$

Equilibrium law equation: $K = \dfrac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]}$

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1. (a) $SiH_2(g) + 2\ Cl_2(g) \rightleftharpoons SiCl_4(g) + H_2(g)$

Products: $SiCl_4(g)$; $H_2(g)$

Reactants: $SiH_2(g)$; 2 $Cl_2(g)$

Equilibrium law equation: $K = \dfrac{[SiCl_4(g)][H_2(g)]}{[SiH_2(g)][Cl_2(g)]^2}$

(b) $2\ PBr_3(g) + 3\ Cl_2(g) \rightleftharpoons 2\ PCl_3(g) + 3\ Br_2(g)$

Products: 2 $PCl_3(g)$; 3 $Br_2(g)$

Reactants: 2 $PBr_3(g)$; 3 $Cl_2(g)$

Equilibrium law equation: $K = \dfrac{[PCl_3(g)]^2[Br_2(g)]^3}{[PBr_3(g)]^2[Cl_2(g)]^3}$

(c) $H_2O(l) \rightleftharpoons H_2O(g)$
Product: $H_2O(g)$
Reactant: $H_2O(l)$
Equilibrium law equation: $K = [H_2O(g)]$

(d) $NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$
Products: $Na_2CO_3(s)$; $CO_2(g)$; $H_2O(g)$
Reactant: 2 $NaHCO_3(s)$
Equilibrium law equation: $K = [CO_2(g)][H_2O(g)]$

2. Given: $2\ NOCl(g) \rightleftharpoons 2\ NO(g) + Cl_2(g)$; volume = 3.0 L;
quantities at equilibrium: $Cl_2(g)$ = 2.4 mol; NOCl(g) = 1.0 mol; NO(g) = 4.5×10^{-3} mol
Required: K
Solution:
Step 1. Write the equilibrium law equation using the balanced chemical equation.

$$K = \frac{[NO(g)]^2[Cl_2(g)]}{[NOCl(g)]^2}$$

Step 2. Calculate the equilibrium concentrations of the reactant and products.

$$[Cl_2(g)]_{equilibrium} = \frac{2.4\ mol}{3.0\ L}$$

$$[Cl_2(g)]_{equilibrium} = 0.8\ mol/L$$

$$[NOCl(g)]_{equilibrium} = \frac{1.0\ mol}{3.0\ L}$$

$$[NOCl(g)]_{equilibrium} = 0.33\ mol/L$$

$$[NO(g)]_{equilibrium} = \frac{4.5 \times 10^{-3}\ mol}{3.0\ L}$$

$$[NO(g)]_{equilibrium} = 1.5 \times 10^{-3}\ mol/L$$

Step 3. Substitute the equilibrium concentrations into the equilibrium law equation and solve for K.

$$K = \frac{[NO(g)]^2[Cl_2(g)]}{[NOCl(g)]^2}$$

$$= \frac{(1.5 \times 10^{-3})^2(0.80)}{(0.33)^2}$$

$$K = 1.6 \times 10^{-5}$$

Statement: The equilibrium constant, K, for the reaction is 1.6×10^{-5}.

3. Given: $4\ Fe(s) + 3\ O_2(g) \rightleftharpoons 2\ Fe_2O_3(s)$; volume = 2.0 L;

quantities at equilibrium: Fe(s) = 1.0 mol; $O_2(g) = 1.0 \times 10^{-3}$ mol; $Fe_2O_3(s)$ = 2.0 mol

Required: *K*

Solution:

Step 1. Write the equilibrium law equation using the balanced chemical equation.

$$K = \frac{1}{[O_2(g)]^3}$$

Step 2. Calculate the equilibrium concentration of oxygen gas.

$$[O_2(g)]_{\text{equilibrium}} = \frac{1.0 \times 10^{-3}\ \text{mol}}{2.0\ \text{L}}$$

$$[O_2(g)]_{\text{equilibrium}} = 5.0 \times 10^{-4}\ \text{mol/L}$$

Step 3. Substitute the equilibrium concentration of oxygen gas into the equilibrium law equation and solve for *K*.

$$K = \frac{1}{[O_2(g)]^3}$$

$$= \frac{1}{(5.0 \times 10^{-4})^3}$$

$$K = 8.0 \times 10^9$$

Statement: The equilibrium constant, *K,* for the reaction is 8.0×10^9.

4.

Comparison of Homogeneous and Heterogeneous Equilibria

	Homogeneous equilibrium	**Heterogeneous equilibrium**
reactants and products	• all same state of matter (either gases or solutions) *Sample chemical equation:* $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$	• present in at least two different states (e.g., gases and solids) *Sample chemical equation:* $PCl_5(s) \rightleftharpoons PCl_3(l) + Cl_2(g)$
equilibrium law equation	• all reactants and products appear *Sample equilibrium law equation:* $K = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$	• pure solids and liquids are not included *Sample equilibrium law equation:* $K = [Cl_2(g)]$
explanation	• Concentrations of substances in gas state or solution can vary.	• Concentrations of pure substances and liquids cannot change; they are constants.

5. Formation of methanol gas from carbon monoxide gas and hydrogen gas:

(a) Unbalanced chemical equation: $CO(g) + H_2(g) \rightleftharpoons CH_3OH(g)$

Balanced chemical equation: $CO(g) + 2\ H_2(g) \rightleftharpoons CH_3OH(g)$

(b) $K = \dfrac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$

If $K = 6.3 \times 10^{-3}$, the concentration of methanol would be expected to be much lower than the concentration of carbon monoxide and hydrogen.

(c) This equilibrium system would not be useful for generating large quantities of methanol because the equilibrium favours the reactants.

(d) The equilibrium constant for the reverse reaction, K', is the reciprocal of K.

$$\begin{aligned} K' &= \frac{1}{K} \\ &= \frac{1}{6.3 \times 10^{-3}} \\ K' &= 160 \end{aligned}$$

6. Formation of carbon monoxide from coal and carbon dioxide gas:

(a) Unbalanced chemical equation: $C(s) + CO_2(g) \rightleftharpoons CO(g)$

Balanced chemical equation: $C(s) + CO_2(g) \rightleftharpoons 2\ CO(g)$

(b) Product: 2 $CO(g)$

Reactants: $C(s)$; $CO_2(g)$

$K = \dfrac{[CO(g)]^2}{[CO_2(g)]}$

(c) More carbon monoxide gas will be produced at 1000 °C than at 25 ºC because the equilibrium constant is much higher at 1000 °C (1.7×10^2 compared to 1.6×10^{-2}), indicating that equilibrium favours the product more at the higher temperature than at the lower temperature.

(d) Answers may vary. Sample answer: Some advantages of a gaseous fuel over a solid fuel are that gaseous fuels can be delivered through pipelines, so it is easier to control their flow into a combustion chamber and they can disperse throughout the volume so they are likely to burn faster.

(e) Sample answer. Some safety issues involved in working with carbon monoxide gas at high temperatures are that leaks of high-temperature gas can cause severe burns and that carbon monoxide is toxic.