Chapter 18
Study Guide - Answers

1. Write the expression for the equilibrium constant for each of the following reactions.

a. $H_2(g) + Cl_2(g) \leftrightarrows 2HCl(g)$

$$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]}$$

b. $N_2(g) + 3\ H_2(g) \leftrightarrows 2\ NH_3(g)$

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

c. $HCN(g) \leftrightarrows H^+(aq) + CN^-(aq)$

$$K_{eq} = \frac{[H^+][CN^-]}{[HCN]}$$

d. $H_2SO_4(g) \leftrightarrows H^+(aq) + HSO^{4-}(aq)$

$$K_{eq} = \frac{[H^+][HSO^{4-}]}{[H_2SO_4]}$$

e. $PbF_2(s) \leftrightarrows Pb^{2+}(aq) + 2F^-(aq)$

$$K_{eq} = [Pb^{2+}][F^-]^2$$

2. From the data provided below, calculate the value of the equilibrium constant for the reaction.

a. $H_2(g) + Cl_2(g) \leftrightarrows 2HCl(g)$

$[H_2] = [Cl_2] = 1.0 \times 10^{-2}$; $[HCl] = 1.0 \times 10^{-4}$

$$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]}$$

$$K_{eq} = \frac{[1.0 \times 10^{-4}]^2}{[1.0 \times 10^{-2}][\ 1.0 \times 10^{-2}]} = 0.0001$$

b. $N_2(g) + 3\ H_2(g) \leftrightarrows 2\ NH_3(g)$

$[N_2] = 4.4 \times 10^{-2}$; $[H_2] = 1.2 \times 10^{-1}$; $[NH_3] = 3.4 \times 10^{-3}$

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_{eq} = \frac{[3.4 \times 10^{-3}]^2}{[4.4 \times 10^{-2}][\ 1.2 \times 10^{-1}]^3} = 0.152$$

3. For the reaction $N_2O_4\,(g) \leftrightarrows 2NO_2\,(g)$, $K_{eq} = 0.2$. At a particular time, the following concentrations are measured; $[N_2O_4] = 2.0\ M$, $[NO_2] = 0.2\ M$. Is this reaction at equilibrium? If not, in which direction will the reaction proceed?

$$N_2O_4\,(g) \leftrightarrows 2NO_2\,(g)$$

$[N_2O_4] = 2.0\ M$
$[NO_2] = 0.2\ M$

$$Q = \frac{[NO_2]^2}{[N_2O_4]}$$

$$Q = \frac{[0.2]^2}{[2.0]} = 0.02$$

$K_{eq} = 0.2$

$Q < K_{eq}$, so the reaction will proceed to the right (forward).

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4. For the reaction 2ICl (g) ⇆ I_2 (g) + Cl_2 (g), K_{eq} = 0.11. At a particular time, the following concentration are measured; [ICl] = 2.5*M*, [I_2] = 2.0*M*, [Cl_2] = 1.2*M*. Is this reaction at equilibrium? If not, in which direction will the reaction proceed?

$$2ICl\ (g) \leftrightarrows I_2\ (g) + Cl_2\ (g)$$

[ICl] = 2.5 *M*
[I_2] = 2.0 *M*
[Cl_2] = 1.2 *M*

$$Q = \frac{[I_2][Cl_2]}{[ICl]^2}$$

$$Q = \frac{[2.0][1.2]}{[2.5]^2} = 0.384$$

K_{eq} = 0.11

Q > K_{eq}, so the reaction will proceed to the left (reverse).

5. Name the 3 stresses that may be applied to a chemical equilibrium to cause it to shift.
   a. Concentration
   b. Temperature
   c. Pressure (gases only)

6. Describe LeChatelier's Principle.
   When a system at equilibrium is subjected to a stress (a change in concentration, temperature, or pressure), the equilibrium will shift in the direction that tends to counteract the effect of the stress.

7. For the reaction below, mark whether the stress listed will cause the reaction to move forward or in reverse.

$$Zn(s) + 2HCl(l) \leftrightarrows ZnCl_2(l) + H_2(g);\ \Delta H = -235kJ$$

a.	Increase Heat	e.	Increase [$ZnCl_2$]	i.	Decrease [H_2]
b.	Increase Pressure	f.	Increase [HCl]	j.	Decrease Pressure
c.	Increase [H_2]	g.	Decrease [HCl]	k.	Decrease Heat
d.	Increase [Zn]	h.	Decrease [$ZnCl_2$]		

   a. Forward reaction is exothermic (heat is a product), so increase in heat drives reaction in reverse.
   b. Reverse reaction requires less volume (no gases produced), so reaction driven in reverse.
   c. Reverse to consume added H_2.
   d. Forward to consume Zn.
   e. Reverse to consume $ZnCl_2$.
   f. Forward to consume HCl.
   g. Reverse to replace HCl.
   h. Forward to replace $ZnCl_2$.
   i. Forward to replace H_2.
   j. Forward.
   k. Forward.

8. In a system at equilibrium, the rates of the forward & reverse reactions are equal.

9. In a system at equilibrium, the concentrations of the forward & reverse reactions are constant.

10. K_{sp} depends on the temperature of the solution.

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11. How many ions are produced by the dissolution of a unit of $CaCl_2$?
Three ions are produced for each unit of $CaCl_2$ – 1 Ca^{2+} and 2 Cl^-.

12. When referred to in terms of K_{sp}, Q is called the ion product. If $Q > K_{sp}$, the solution is supersaturated and a precipitate will form. If $Q < K_{sp}$, the solution is unsaturated and more solute may dissolve. If $Q = K_{sp}$, the solution is at equilibrium.

13. What is the common-ion effect?
Shift in equilibrium that occurs because the concentration of an ion that is part of the equilibrium is changed. For example, if you dissolve $MgCl_2$ in water, this is what happens:

$$MgCl_2(s) \leftrightarrow Mg^{2+}(aq) + 2Cl^-(aq)$$

Ions of both Mg^{2+} and Cl^- would be floating around in the water. If you then add a second solution that also has NaCl [$NaCl(s) \leftrightarrow Na^+(aq) + Cl^-(aq)$ possibly], the added Cl^- would cause the above reaction to shift in reverse and $MgCl_2(s)$ would precipitate out of the solution. The "common ion" affects the equilibrium of one of the reactions (LeChatelier's Principle).

14. Write the expression for the solubility product constant for $SrSO_4$.

$$SrSO_4(s) \leftrightarrow Sr^{2+}(aq) + SO_4^{2-}(aq)$$
$$K_{sp} = [Sr^{2+}][SO_4^{2-}]$$

15. Write the expression for the solubility product constant for $A1_2(SO_4)_3$.

$$Al_2(SO4)_3(s) \leftrightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$$
$$K_{sp} = [Al^{3+}]^2[SO_4^{2-}]^3$$

16. Write the expression for the solubility product constant for AgI.

$$AgI(s) \leftrightarrow Ag^+(aq) + I^-(aq)$$
$$K_{sp} = [Ag^+][I^-]$$

17. A sample of $SrCO_3(s)$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Sr^{2+} is $4.0 \times 10^{-5} M$ at equilibrium. What is the value of K_{sp} for $SrCO_3$?

$$SrCO_3(s) \leftrightarrow Sr^{2+}(aq) + CO_3^{2-}(aq)$$
$$K_{sp} = [Sr^{2+}][CO_3^{2-}]$$
The concentrations of Sr^{2+} and CO_3^{2-} will be equal because their mole ratio is 1:1
$$K_{sp} = [4.0 \times 10^{-5}][4.0 \times 10^{-5}]$$
$$Ksp = 1.6 \times 10^{-9}$$

18. A sample of $BaSO_4(s)$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Ba^{2+} is $1.05 \times 10^{-5} M$ at equilibrium. What is the value of K_{sp} for $BaSO_4$?

$$BaSO_4(s) \leftrightarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$$
$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$
The concentrations of Ba^{2+} and SO_4^{2-} will be equal because their mole ratio is 1:1
$$K_{sp} = [1.05 \times 10^{-5}][1.05 \times 10^{-5}]$$
$$Ksp = 1.10 \times 10^{-10}$$

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Added question:

A sample of $AlCl_3(s)$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Al^{3+} is $3.05 \times 10^{-3} M$ at equilibrium. What is the value of K_{sp} for $AlCl_3$?

$$AlCl_3(s) \leftrightarrow Al^{3+}(aq) + 3Cl^-(aq)$$

$$K_{sp} = [Al^{3+}][Cl^-]^3$$

The concentration of Cl^- will be 3x the concentration of Al^{3+} because the mole ratio of Cl^- to Al^{3+} is 3:1

$$K_{sp} = [3.05 \times 10^{-3}][9.15 \times 10^{-3}]^3$$

$$K_{sp} = 2.34 \times 10^{-9}$$

19. What will be the equilibrium concentration of dissolved ions in a saturated solution of $Pb(OH)_2$ at 25°C? K_{sp} for the reaction is 1.2×10^{-15}.

$$Pb(OH)_2(s) \leftrightarrow Pb^{2+}(aq) + 2OH^-(aq)$$

$$K_{sp} = [Pb^{2+}][OH^-]^2$$

Setting the concentration of Pb^{2+} to x, and the concentration of OH^- to $2x$, (the mole ratio of Pb^{2+} to OH^- is 1:2), we get the following:

$$1.2 \times 10^{-15} = [x][2x]^2$$

$$1.2 \times 10^{-15} = 4x^3$$

$$3.0 \times 10^{-16} = x^3$$

$$6.69 \times 10^{-6} = x$$

So, $[Pb^{2+}]$ is $6.69 \times 10^{-6} M$ and $[OH^-]$ would be twice that or $1.34 \times 10^{-5} M$.

20. What will be the equilibrium concentration of dissolved ions in a saturated solution of $SrSO_4$ at 25°C? K_{sp} for the reaction is 3.44×10^{-7}.

$$SrSO_4(s) \leftrightarrow Sr^{2+}(aq) + SO_4^{2-}(aq)$$

$$K_{sp} = [Sr^{2+}][SO_4^{2-}]$$

Setting the concentration of Sr^{2+} to x, and the concentration of SO_4^{2-} also to x, (the mole ratio of Sr^{2+} to SO_4^{2-} is 1:1), we get the following:

$$3.44 \times 10^{-7} = [x][x]$$

$$3.44 \times 10^{-7} = x^2$$

$$5.87 \times 10^{-4} = x$$

So, $[Sr^{2+}]$ is $5.87 \times 10^{-4} M$ and $[SO_4^{2-}]$ would be the same.