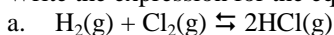


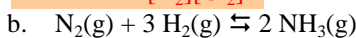
# Chapter 18

## Study Guide - Answers

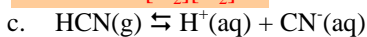
1. Write the expression for the equilibrium constant for each of the following reactions.



$$K_{\text{eq}} = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]}$$



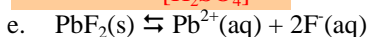
$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$



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

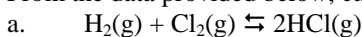


$$K_{\text{eq}} = \frac{[\text{H}^+][\text{HSO}_4^-]}{[\text{H}_2\text{SO}_4]}$$



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

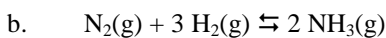
2. From the data provided below, calculate the value of the equilibrium constant for the reaction.



$$[\text{H}_2] = [\text{Cl}_2] = 1.0 \times 10^{-2}; [\text{HCl}] = 1.0 \times 10^{-4}$$

$$K_{\text{eq}} = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]}$$

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



$$[\text{N}_2] = 4.4 \times 10^{-2}; [\text{H}_2] = 1.2 \times 10^{-1}; [\text{NH}_3] = 3.4 \times 10^{-3}$$

$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

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

3. For the reaction  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ ,  $K_{\text{eq}} = 0.2$ . At a particular time, the following concentrations are measured;  $[\text{N}_2\text{O}_4] = 2.0 \text{ M}$ ,  $[\text{NO}_2] = 0.2 \text{ M}$ . Is this reaction at equilibrium? If not, in which direction will the reaction proceed?



$$[\text{N}_2\text{O}_4] = 2.0 \text{ M}$$

$$[\text{NO}_2] = 0.2 \text{ M}$$

$$Q = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

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

$$K_{\text{eq}} = 0.2$$

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

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4. For the reaction  $2\text{ICl (g)} \rightleftharpoons \text{I}_2\text{(g)} + \text{Cl}_2\text{(g)}$ ,  $K_{\text{eq}} = 0.11$ . At a particular time, the following concentration are measured;  $[\text{ICl}] = 2.5\text{M}$ ,  $[\text{I}_2] = 2.0\text{M}$ ,  $[\text{Cl}_2] = 1.2\text{M}$ . Is this reaction at equilibrium? If not, in which direction will the reaction proceed?



$$[\text{ICl}] = 2.5\text{ M}$$

$$[\text{I}_2] = 2.0\text{ M}$$

$$[\text{Cl}_2] = 1.2\text{ M}$$

$$Q = \frac{[\text{I}_2][\text{Cl}_2]}{[\text{ICl}]^2}$$

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

$$K_{\text{eq}} = 0.11$$

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

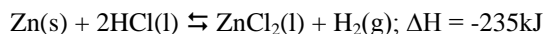
5. Name the 3 stresses that may be applied to a chemical equilibrium to cause it to shift.

- Concentration
- Temperature
- 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.



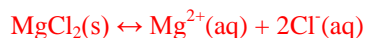
- |                            |                               |                            |
|----------------------------|-------------------------------|----------------------------|
| a. Increase Heat           | e. Increase $[\text{ZnCl}_2]$ | i. Decrease $[\text{H}_2]$ |
| b. Increase Pressure       | f. Increase $[\text{HCl}]$    | j. Decrease Pressure       |
| c. Increase $[\text{H}_2]$ | g. Decrease $[\text{HCl}]$    | k. Decrease Heat           |
| d. Increase $[\text{Zn}]$  | h. Decrease $[\text{ZnCl}_2]$ |                            |
- Forward reaction is exothermic (heat is a product), so increase in heat drives reaction in reverse.
  - Reverse reaction requires less volume (no gases produced), so reaction driven in reverse.
  - Reverse to consume added  $\text{H}_2$ .
  - Forward to consume  $\text{Zn}$ .
  - Reverse to consume  $\text{ZnCl}_2$ .
  - Forward to consume  $\text{HCl}$ .
  - Reverse to replace  $\text{HCl}$ .
  - Forward to replace  $\text{ZnCl}_2$ .
  - Forward to replace  $\text{H}_2$ .
  - Forward.
  - 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.  $K_{\text{sp}}$  depends on the temperature of the solution.

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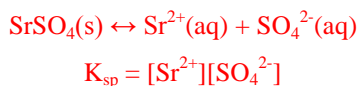
## Study Guide - Answers

11. How many ions are produced by the dissolution of a unit of  $\text{CaCl}_2$ ?  
Three ions are produced for each unit of  $\text{CaCl}_2$  – 1  $\text{Ca}^{2+}$  and 2  $\text{Cl}^-$ .
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  $\text{MgCl}_2$  in water, this is what happens:

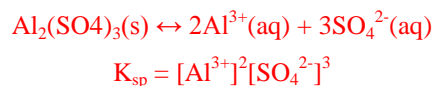


Ions of both  $\text{Mg}^{2+}$  and  $\text{Cl}^-$  would be floating around in the water. If you then add a second solution that also has  $\text{NaCl}$  [ $\text{NaCl}(\text{s}) \leftrightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$  possibly], the added  $\text{Cl}^-$  would cause the above reaction to shift in reverse and  $\text{MgCl}_2(\text{s})$  would precipitate out of the solution. The “common ion” affects the equilibrium of one of the reactions (LeChatelier’s Principle).

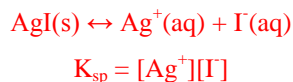
14. Write the expression for the solubility product constant for  $\text{SrSO}_4$ .



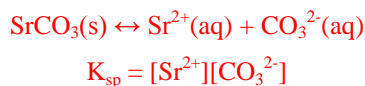
15. Write the expression for the solubility product constant for  $\text{Al}_2(\text{SO}_4)_3$ .



16. Write the expression for the solubility product constant for  $\text{AgI}$ .



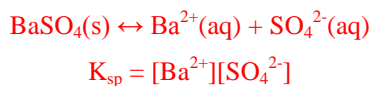
17. A sample of  $\text{SrCO}_3(\text{s})$  is added to pure water and allowed to come to equilibrium at  $25^\circ\text{C}$ . The concentration of  $\text{Sr}^{2+}$  is  $4.0 \times 10^{-5} \text{ M}$  at equilibrium. What is the value of  $K_{sp}$  for  $\text{SrCO}_3$ ?



The concentrations of  $\text{Sr}^{2+}$  and  $\text{CO}_3^{2-}$  will be equal because their mole ratio is 1:1

$$K_{sp} = [4.0 \times 10^{-5}][4.0 \times 10^{-5}]$$
$$K_{sp} = 1.6 \times 10^{-9}$$

18. A sample of  $\text{BaSO}_4(\text{s})$  is added to pure water and allowed to come to equilibrium at  $25^\circ\text{C}$ . The concentration of  $\text{Ba}^{2+}$  is  $1.05 \times 10^{-5} \text{ M}$  at equilibrium. What is the value of  $K_{sp}$  for  $\text{BaSO}_4$ ?



The concentrations of  $\text{Ba}^{2+}$  and  $\text{SO}_4^{2-}$  will be equal because their mole ratio is 1:1

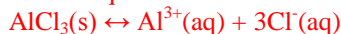
$$K_{sp} = [1.05 \times 10^{-5}][1.05 \times 10^{-5}]$$
$$K_{sp} = 1.10 \times 10^{-10}$$

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Added question:

A sample of  $\text{AlCl}_3(\text{s})$  is added to pure water and allowed to come to equilibrium at  $25^\circ\text{C}$ . The concentration of  $\text{Al}^{3+}$  is  $3.05 \times 10^{-3} \text{ M}$  at equilibrium. What is the value of  $K_{\text{sp}}$  for  $\text{AlCl}_3$ ?



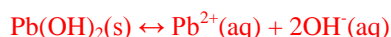
$$K_{\text{sp}} = [\text{Al}^{3+}][\text{Cl}^{-}]^3$$

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

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

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

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



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

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

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

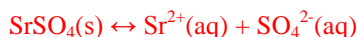
$$1.2 \times 10^{-15} = 4x^3$$

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

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

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

20. What will be the equilibrium concentration of dissolved ions in a saturated solution of  $\text{SrSO}_4$  at  $25^\circ\text{C}$ ?  $K_{\text{sp}}$  for the reaction is  $3.44 \times 10^{-7}$ .



$$K_{\text{sp}} = [\text{Sr}^{2+}][\text{SO}_4^{2-}]$$

Setting the concentration of  $\text{Sr}^{2+}$  to  $x$ , and the concentration of  $\text{SO}_4^{2-}$  also to  $x$ , (the mole ratio of  $\text{Sr}^{2+}$  to  $\text{SO}_4^{2-}$  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,  $[\text{Sr}^{2+}]$  is  $5.87 \times 10^{-4} \text{ M}$  and  $[\text{SO}_4^{2-}]$  would be the same.