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Skills Worksheet

Problem Solving

Equilibrium

Not all processes in nature proceed to completion. In fact, most changes hover somewhere between the initial state and what would be the final state. Compare a light switch and a dimmer. If the mechanical switch is working properly, it can be stable in only two positions: on or off. Either current flows or it doesn't. With a dimmer you can regulate the flow of current so that it stays somewhere between fully on and fully off. If you've used mechanical balances, you know that to weigh an object accurately you must adjust the masses so that the pointer hovers in the middle of its range. The balance is in a state of equilibrium.

Most chemical reactions also reach a state of equilibrium between no reaction at all and the complete reaction to form products. Equilibrium states occur when reactions are reversible, that is, when products react to re-form the original reactants. When the products re-form reactants at the same rate as the reactants form the products, then the equilibrium point of the reaction has been reached.

The progress of a reaction is gauged by measuring the concentrations in moles per liter of reactants and products. At the equilibrium point, these concentrations stop changing.

$$A + 2B \rightleftarrows 2C$$

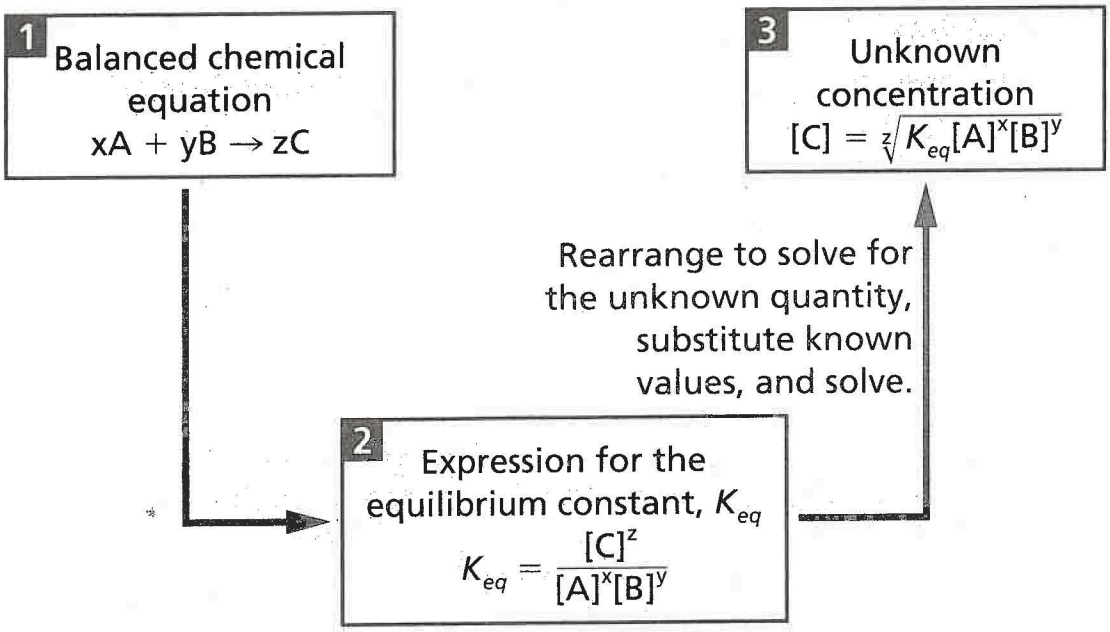
The equation above represents a reaction in which 1 mol of A reacts with 2 mol of B to produce 2 mol of C. As C is formed, it breaks down to re-form reactants. The extent of the reaction at equilibrium is indicated by the equilibrium constant, K_{eq}.

$$K_{eq} = \frac{[C]^2}{[A][B]}$$

As you can see, the concentration of each reaction component is raised to the power of its coefficient in the balanced equation. These concentration terms are arranged in a fraction with products in the numerator and reactants in the denominator. Pure substances (substances that appear in the chemical equation as solids or pure liquids) are not included in the equilibrium expressions because their concentrations are meaningless.

Problems involving chemical equilibrium will ask you to solve for either the equilibrium constant, K_{eq}, given the concentrations of all of the reaction components, or the concentration of one of the reaction components, given K_{eq}.

General Plan for Solving Equilibrium Problems



Sample Problem 1

The following equation represents the reversible decomposition of PCl_5.

$$PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$$

At 250°C, the equilibrium concentrations of the substances are as follows:

$[PCl_5] = 1.271$ M

$[PCl_3] = 0.229$ M

$[Cl_2] = 0.229$ M

What is the value of the equilibrium constant, K_{eq}, for this reaction?

Solution

ANALYZE

What is given in the problem? **the equilibrium concentrations of the products and reactant at 250°C**

What are you asked to find? **the value of the equilibrium constant for the reaction**

Items	Data
Molar concentration of PCl_5 at equilibrium	1.271
Molar concentration of PCl_3 at equilibrium	0.229 M
Molar concentration of Cl_2 at equilibrium	0.229 M
Equilibrium constant K_{eq}	?

PLAN

What steps are needed to calculate the equilibrium constant for the given reaction?

Set up the equilibrium expression for the reaction using the coefficients as exponents. Substitute the concentration values, and calculate K_{eq}.

$$PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g) \xrightarrow{\text{1}} K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]} \quad \text{2}$$

Note that since all coefficients have the value 1, there is no need to write in the exponent.

$$K_{eq} = \frac{\overset{given}{[PCl_3]}\overset{given}{[Cl_2]}}{\underset{given}{[PCl_5]}}$$

COMPUTE

$$K_{eq} = \frac{[0.229][0.229]}{[1.271]} = 0.0413$$

EVALUATE

Are the units correct?

Yes; the equilibrium constant has no units.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because data values were given to a minimum of three significant figures.

Is the answer reasonable?

Yes; the calculation may be approximated as (0.2 × 0.2)/1. This approximation gives a result of 0.04, which is close to the calculated value.

Practice

1. Calculate the equilibrium constants for the following hypothetical reactions. Assume that all components of the reactions are gaseous.

   a. $A \rightleftarrows C + D$

   At equilibrium, the concentration of A is 2.24×10^{-2} M and the concentrations of both C and D are 6.41×10^{-3} M. **ans: $K_{eq} = 1.83 \times 10^{-3}$**

b. $A + B \rightleftarrows C + D$

At equilibrium, the concentrations of both A and B are 3.23×10^{-5} M and the concentrations of both C and D are 1.27×10^{-2} M. **ans: $K_{eq} = 1.55 \times 10^{5}$**

c. $A + B \rightleftarrows 2C$

At equilibrium, the concentrations of both A and B are 7.02×10^{-3} M and the concentration of C is 2.16×10^{-2} M. **ans: $K_{eq} = 9.47$**

d. $2A \rightleftarrows 2C + D$

At equilibrium, the concentration of A is 6.59×10^{-4} M. The concentration of C is 4.06×10^{-3} M, and the concentration of D is 2.03×10^{-3} M.
ans: $K_{eq} = 7.71 \times 10^{-2}$

e. $A + B \rightleftarrows C + D + E$

At equilibrium, the concentrations of both A and B are 3.73×10^{-4} M and the concentrations of C, D, and E are 9.35×10^{-4} M. **ans: $K_{eq} = 5.88 \times 10^{-3}$**

f. $2A + B \rightleftarrows 2C$

At equilibrium, the concentration of A is 5.50×10^{-3} M, the concentration of B is 2.25×10^{-3} M, and the concentration of C is 1.02×10^{-2} M.
ans: $K_{eq} = 1.53 \times 10^{3}$

Sample Problem 2

The following equilibrium reaction is used in the manufacture of methanol. The equilibrium constant at 400 K for the reaction is 1.609.

$$CO(g) + 2H_2(g) \rightleftarrows CH_3OH(g)$$

At equilibrium, the mixture in the reaction vessel has a concentration of 0.818 M of CH_3OH and 1.402 M of CO. Calculate the concentration of H_2 in the vessel.

Solution

ANALYZE

What is given in the problem? **the equilibrium concentrations of CO and CH_3OH, and the equilibrium constant at 400 K**

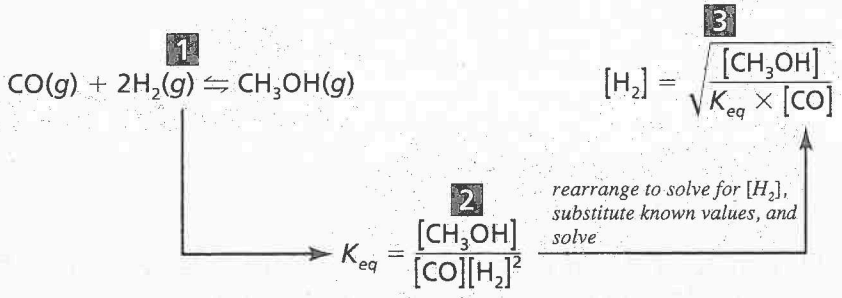
What are you asked to find? **the equilibrium concentration of H_2 in the vessel**

Items	Data
Molar concentration of CO at equilibrium	1.402 M
Molar concentration of H_2 at equilibrium	? M
Molar concentration of CH_3OH at equilibrium	0.818 M
Equilibrium constant K_{eq}	0.609 M

PLAN

What steps are needed to calculate the concentration of H_2?

Set up the equilibrium expression for the reaction using coefficients as exponents. Rearrange the expression to solve for $[H_2]$. Substitute known values for $[CO_2]$, $[CH_3OH]$, and K_{eq}, and solve for $[H_2]$.



COMPUTE

$$[H_2] = \sqrt{\frac{[0.818]}{1.609 \times [1.402]}} = 0.602 \text{ M}$$

EVALUATE

Are the units correct?
Yes; concentrations are in moles per liter.

Is the number of significant figures correct?
Yes; the number of significant figures is correct because data values were given to a minimum of three significant figures.

Is the answer reasonable?
Yes; the calculation may be approximated as $[1 \div (1.5 \times 1.5)]^{1/2} = 0.67$, which is close to the calculated value.

Practice

1. Calculate the concentration of product D in the following hypothetical reaction:

$$2A(g) \rightleftarrows 2C(g) + D(g)$$

At equilibrium, the concentration of A is 1.88×10^{-1} M, the concentration of C is 6.56 M, and the equilibrium constant is 2.403×10^{2}. **ans: 0.197 M**

2. At a temperature of 700 K, the equilibrium constant is 3.164×10^{3} for the following reaction system for the hydrogenation of ethene, C_2H_4, to ethane, C_2H_6.

$$C_2H_4(g) + H_2(g) \rightleftarrows C_2H_6(g)$$

What will be the equilibrium concentration of ethene if the concentration of H_2 is 0.0619 M and the concentration of C_2H_6 is 1.055 M? **ans: 5.39×10^{-3} M**

Additional Problems—Equilibrium

1. Using the reaction $A + 2B \rightleftarrows C + 2D$, determine K_{eq} if the following equilibrium concentrations are found. All components are gases.

$[A] = 0.0567\ M$

$[B] = 0.1171\ M$

$[C] = 0.000\ 3378\ M$

$[D] = 0.000\ 6756\ M$

2. In the reaction $2A \rightleftarrows 2C + 2D$, determine K_{eq} when the following equilibrium concentrations are found. All components are gases.

$[A] = 0.1077\ M$

$[C] = 0.000\ 4104\ M$

$[D] = 0.000\ 4104\ M$

3. Calculate the equilibrium constant for the following reaction. Note the phases of the components.

$$2A(g) + B(s) \rightleftarrows C(g) + D(g)$$

The equilibrium concentrations of the components are

$[A] = 0.0922\ M$

$[C] = 4.11 \times 10^{-4}\ M$

$[D] = 8.22 \times 10^{-4}\ M$

4. The equilibrium constant of the following reaction for the decomposition of phosgene at 25°C is 4.282×10^{-2}.

$$COCl_2(g) \rightleftarrows CO(g) + Cl_2(g)$$

a. What is the concentration of $COCl_2$ when the concentrations of both CO and Cl_2 are $5.90 \times 10^{-3}\ M$?

b. When the equilibrium concentration of $COCl_2$ is 0.003 70 M, what are the concentrations of CO and Cl_2? Assume the concentrations are equal.

5. Consider the following hypothetical reaction.

$$A(g) + B(s) \rightleftarrows C(g) + D(s)$$

a. If $K_{eq} = 1$ for this reaction at 500 K, what can you say about the concentrations of A and C at equilibrium?

b. If raising the temperature of the reaction results in an equilibrium with a higher concentration of C than A, how will the value of K_{eq} change?

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Problem Solving *continued*

6. The following reaction occurs when steam is passed over hot carbon. The mixture of gases it generates is called water gas and is useful as an industrial fuel and as a source of hydrogen for the production of ammonia

$$C(s) + H_2O(g) \rightleftarrows CO(g) + H_2(g)$$

The equilibrium constant for this reaction is 4.251×10^{-2} at 800 K. If the equilibrium concentration of $H_2O(g)$ is 0.1990 M, what concentrations of CO and H_2 would you expect to find?

7. When nitrogen monoxide gas comes in contact with air, it oxidizes to the brown gas nitrogen dioxide according to the following equation:

$$2NO(g) + O_2(g) \rightleftarrows 2NO_2(g)$$

a. The equilibrium constant for this reaction at 500 K is 1.671×10^{4}. What concentration of NO_2 is present at equilibrium if $[NO] = 6.200 \times 10^{-2}$ M and $[O_2] = 8.305 \times 10^{-3}$ M?

b. At 1000 K, the equilibrium constant, K_{eq}, for the same reaction is 1.315×10^{-2}. What will be the concentration of NO_2 at 1000 K given the same concentrations of NO and O_2 as were in (a)?

8. Consider the following hypothetical reaction, for which $K_{eq} = 1$ at 300 K:

$$A(g) + B(g) \rightleftarrows 2C(g)$$

a. If the reaction begins with equal concentrations of A and B and a zero concentration of C, what can you say about the relative concentrations of the components at equilibrium?

b. Additional C is introduced at equilibrium, and the temperature remains constant. When equilibrium is restored, how will the concentrations of all components have changed? How will K_{eq} have changed?

9. The equilibrium constant for the following reaction of hydrogen gas and bromine gas at 25°C is 5.628×10^{18}.

$$H_2(g) + Br_2(g) \rightleftarrows 2HBr(g)$$

a. Write the equilibrium expression for this reaction.

b. Assume that equimolar amounts of H_2 and Br_2 were present at the beginning. Calculate the equilibrium concentration of H_2 if the concentration of HBr is 0.500 M.

c. If equal amounts of H_2 and Br_2 react, which reaction component will be present in the greatest concentration at equilibrium? Explain your reasoning.

Problem Solving *continued*

10. The following reaction reaches an equilibrium state:

$$N_2F_4(g) \rightleftarrows 2NF_2(g)$$

At equilibrium at 25°C the concentration of N_2F_4 is found to be 0.9989 M and the concentration of NF_2 is 1.131×10^{-3} M. Calculate the equilibrium constant of the reaction.

11. The equilibrium between dinitrogen tetroxide and nitrogen dioxide is represented by the following equation:

$$N_2O_4(g) \rightleftarrows NO_2(g)$$

A student places a mixture of the two gases into a closed gas tube and allows the reaction to reach equilibrium at 25°C. At equilibrium, the concentration of N_2O_4 is found to be 5.95×10^{-1} M and the concentration of NO_2 is found to be 5.24×10^{-2} M. What is the equilibrium constant of the reaction?

12. Consider the following equilibrium system:

$$NaCN(s) + HCl(g) \rightleftarrows HCN(g) + NaCl(s)$$

a. Write a complete expression for the equilibrium constant of this system.

b. The K_{eq} for this reaction is 2.405×10^6. What is the concentration of HCl remaining when the concentration of HCN is 0.8959 M?

13. The following reaction is used in the industrial production of hydrogen gas:

$$CH_4(g) + H_2O(g) \rightleftarrows CO(g) + 3H_2(g)$$

The equilibrium constant of this reaction at 298 K (25°C) is 3.896×10^{-27}, but at 1100 K the constant is 3.112×10^2.

a. What do these equilibrium constants tell you about the progress of the reaction at the two temperatures?

b. Suppose the reaction mixture is sampled at 1100 K and found to contain 1.56 M of hydrogen, 3.70×10^{-2} M of methane, and 8.27×10^{-1} M of gaseous H_2O. What concentration of carbon monoxide would you expect to find?

14. Dinitrogen tetroxide, N_2O_4, is soluble in cyclohexane, a common nonpolar solvent. While in solution, N_2O_4 can break down into NO_2 according to the following equation:

$$N_2O_4(cyclohexane) \rightleftarrows NO_2(cyclohexane)$$

At 20°C, the following concentrations were observed for this equilibrium reaction:

$[N_2O_4] = 2.55 \times 10^{-3}$ M

$[NO_2] = 10.4 \times 10^{-3}$ M

What is the value of the equilibrium constant for this reaction? Note, the chemical equation must be balanced first.

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Problem Solving *continued*

15. The reaction given in item 14 also occurs when the dinitrogen tetroxide and nitrogen dioxide are dissolved in carbon tetrachloride, CCl_4, another nonpolar solvent.

$$N_2O_4(CCl_4) \rightleftarrows NO_2(CCl_4)$$

The following experimental data were obtained at 20°C:

$[N_2O_4] = 2.67 \times 10^{-3}$ M

$[NO_2] = 10.2 \times 10^{-3}$ M

Calculate the value of the equilibrium constant for this reaction occurring in carbon tetrachloride.

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Skills Worksheet

Problem Solving

Equilibrium of Salts, K_{sp}

When you try to dissolve a solid substance in water, you expect the solid form to disappear, forming ions or molecules in solution. Most substances, however, are only slightly soluble in water. For example, when you stir silver chloride in water, you may think none of the solid dissolves. Does this mean that some of the solid dissolves, forms a saturated solution, and, after that, experiences no further change between the solid and solution phases? It is true to say that there is no further change in the amount of substance in either the solution or the solid phase, but to say, that no further change occurs is inaccurate. An equilibrium exists between the silver chloride and its dissolved ions, and a state of equilibrium is a dynamic state. The following chemical equation shows this equilibrium.

$$AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$$

Like the other examples of equilibria that you have studied, the extent to which this solubility equilibrium proceeds toward the products (the ions in solution) is indicated by an equilibrium constant. When an equilibrium constant is written for a solubility equilibrium, it is called a solubility product constant and is symbolized as K_{sp}.

In solubility equilibrium problems, the reactants are pure substances, and pure substances are never included in an equilibrium expression. That means that you will not have anything in the denominator in the expression for the solubility product constant. This K_{sp} expression for the silver chloride example can be written as follows.

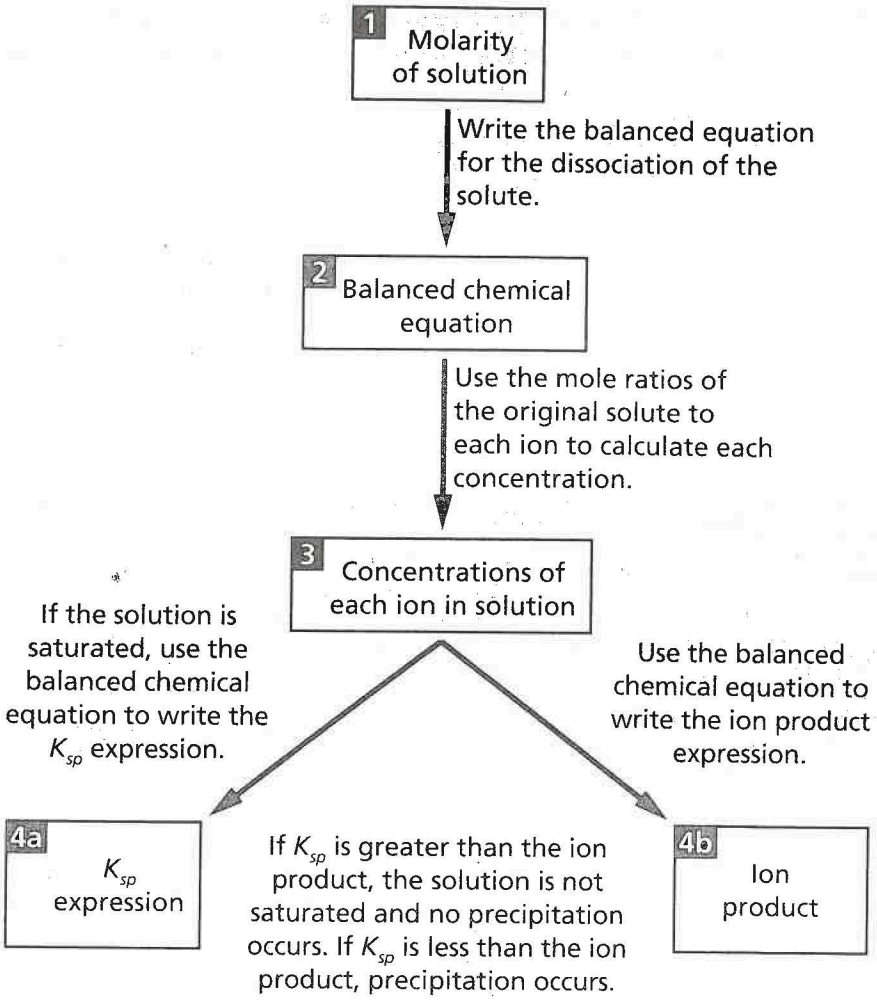
$$K_{sp} = [Ag^+][Cl^-]$$

Note that the coefficients in the balanced equation are understood to be 1 for both silver and chlorine. Therefore, no exponents appear in the K_{sp} expression. As with any equilibrium expression, the concentration of each component is raised to the power of its coefficient from the balanced chemical equation. The value for this solubility product constant is 1.77×10^{-10} at 25°C. The very small value of K_{sp} indicates that silver chloride is only very slightly soluble in aqueous solution at this temperature. The value of K_{sp} supports the observation that little seems to occur when silver chloride is stirred into a water solution.

In this worksheet, you will learn to apply the solubility equilibrium relationship to determine K_{sp} for substances and to calculate concentrations of ions in saturated solutions using their K_{sp} values.

Problem Solving *continued*

General Plan for Solving Solubility Equilibrium Problems



Sample Problem 1

A saturated solution of magnesium fluoride, MgF_2, contains 0.00741 g of dissolved MgF_2 per 1.00×10^2 mL at 25°C. What is the K_{sp} for magnesium fluoride?

Solution

ANALYZE

What is given in the problem? **the mass of MgF_2 dissolved in 1.00×10^2 mL of a saturated solution**

What are you asked to find? **the solubility product constant, K_{sp}**

Items	Data
Mass of dissolved MgF_2	0.00741 g
Volume of solution	1.00×10^2 mL
Molar mass of MgF_2*	62.30 g/mol
Molar concentration of MgF_2	? M
Molar concentration of Mg^{2+}	? M
Molar concentration of F^-	? M
K_{sp} of MgF_2	?

*determined from the periodic table

PLAN

What steps are needed to calculate the solubility product constant, K_{sp}, of MgF_2?
Determine the molar concentration of the saturated MgF_2 solution. Write the balanced chemical equation for the dissociation of MgF_2, and use this equation to determine the concentrations of each ion in solution. Compute K_{sp}.

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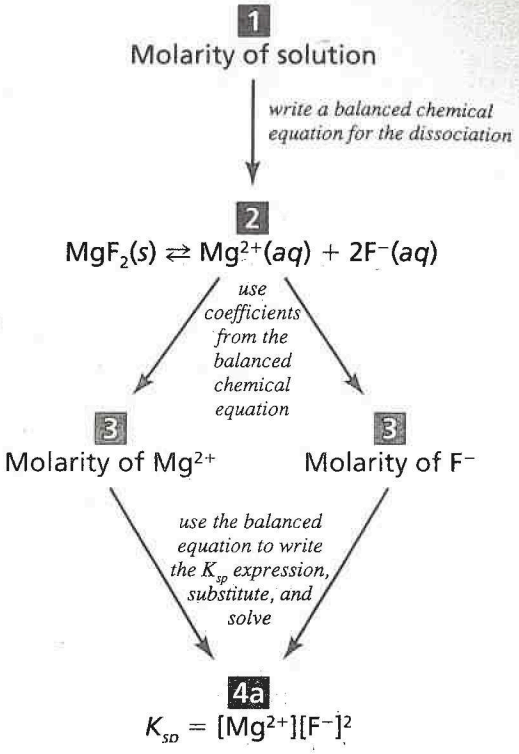
Problem Solving *continued*



Calculate the molarity of the saturated MgF_2 solution.

$$\overset{given}{\text{g MgF}_2} \times \frac{\overset{1}{\overline{\text{molar mass MgF}_2}}}{\frac{1 \text{ mol MgF}_2}{62.30 \text{ g MgF}_2}} = \text{mol MgF}_2$$

$$\overset{given}{\text{mL solution}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = \text{L solution}$$

$$\frac{\overset{calculated\ above}{\text{mol MgF}_2}}{\underset{calculated\ above}{\text{L solution}}} = [MgF_2]$$

Write the balanced chemical equation for the dissociation to determine the mole ratios of solute and ions.

$$MgF_2(s) \overset{a}{} Mg^{2+}(aq) + 2F^-(aq)$$
$$[MgF_2] = [Mg^{2+}]$$
$$2[MgF_2] = [F^-]$$

Write the K_{sp} expression.

$$K_{sp} = [Mg^{2+}][F^-]^2$$

$$K_{sp} = \overset{calculated\ above}{[MgF_2]} (2[MgF_2])^2$$

COMPUTE

$$0.007\,41\ \text{g}\ \cancel{\text{MgF}_2} \times \frac{1\ \text{mol MgF}_2}{62.30\ \text{g}\ \cancel{\text{MgF}_2}} = 1.19 \times 10^{-4}\ \text{mol MgF}_2$$

$$100.\ \cancel{\text{mL}}\ \text{solution} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.100\ \text{L solution}$$

$$\frac{1.19 \times 10^{-4}\ \text{mol MgF}_2}{0.100\ \text{L solution}} = [\text{MgF}_2] = 1.19 \times 10^{-3}\ \text{M}$$

$$[\text{Mg}^{2+}] = [\text{MgF}_2] = 1.19 \times 10^{-3}\ \text{M}$$

$$[\text{F}^-] = 2[\text{MgF}_2] = 2.38 \times 10^{-3}\ \text{M}$$

$$K_{sp} = [1.19 \times 10^{-3}][2.38 \times 10^{-3}]^2 = 6.74 \times 10^{-9}$$

EVALUATE

Are the units correct?

Yes; K_{sp} has no units.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because all data were given to three significant figures.

Is the answer reasonable?

Yes; the calculation can be approximated as $(1 \times 10^{-3})(2.5 \times 10^{-3})^2 = 6 \times 10^{-9}$, which is of the same order of magnitude as the calculated answer.

Practice

1. Silver bromate, $AgBrO_3$, is slightly soluble in water. A saturated solution is found to contain 0.276 g $AgBrO_3$ dissolved in 150.0 mL of water. Calculate K_{sp} for silver bromate. ans: $K_{sp} = 6.09 \times 10^{-5}$

2. 2.50 L of a saturated solution of calcium fluoride leaves a residue of 0.0427 g of CaF_2 when evaporated to dryness. Calculate the K_{sp} of CaF_2.
ans: $K_{sp} = 4.20 \times 10^{-11}$

Sample Problem 2

The K_{sp} for lead(II) iodide is 7.08×10^{-9} at 25°C. What is the molar concentration of PbI_2 in a saturated solution?

Solution

ANALYZE

What is given in the problem — **the solubility product constant, K_{sp} of PbI_2**

What are you asked to find? — **the concentration of PbI_2 in a saturated solution**

Items	Data
Ksp of PbI_2	7.08×10^{-9}
Concentration of Pb^{2+}	?
Concentration of I^-	?
Concentration of PbI_2 in solution	?

PLAN

What steps are needed to calculate the concentration of dissolved PbI_2 in a saturated solution?

Write the equation for the dissociation of PbI_2. Set up the equation for K_{sp}, and compute the concentrations of the ions. Determine the concentration of dissolved solute.

Write the balanced chemical equation for the dissociation of lead(II) iodide, PbI_2 in aqueous solution.

$$PbI_2(s) \rightleftarrows Pb^{2+}(aq) + 2I^-(aq)$$

Write the K_{sp} expression.

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

Substitute x for $[Pb^{2+}]$. The balanced equation gives the following relationship:

$$2[Pb^{2+}] = [I^-].$$

Therefore, $[I^-] = 2x$.

$$K_{sp} = [x][2x]^2$$

Rearrange, and solve for x.

$$K_{sp} = [x][4x^2]$$

$$K_{sp} = 4x^3$$

$$x = \sqrt[3]{\frac{K_{sp}^{given}}{4}} = [Pb^{2+}]$$

Relate the substituted value to the unknown solution concentration using the mole ratio from the original balanced chemical equation. The mole ratio shows that $[Pb^{2+}] = [PbI_2]$.

COMPUTE

$$x = \sqrt[3]{\frac{7.08 \times 10^{-9}}{4}} = [Pb^{2+}] = 1.21 \times 10^{-3}\ M$$

$$[Pb^{2+}] = [PbI_2] = 1.21 \times 10^{-3}\ M$$

EVALUATE

Are the units correct?

Yes; concentrations are in molarity (mol/L).

Is the number of significant figures correct?

Yes; the number of significant figures is correct because data were given to three significant figures.

Is the answer reasonable?

Yes; the best check is to use the result to calculate K_{sp} and see if it gives (or is very near) the K_{sp} you started with. In this case, the calculated K_{sp} is 7.08×10^{-9}, the same value as was given.

Practice

1. The K_{sp} of calcium sulfate, $CaSO_4$, is 9.1×10^{-6}. What is the molar concentration of $CaSO_4$ in a saturated solution? **ans: 3.0×10^{-3} M**

2. A salt has the formula X_2Y, and its K_{sp} is 4.25×10^{-7}.

   a. What is the molarity of a saturated solution of the salt?
   ans: $[X_2Y] = 4.74 \times 10^{-3}$ M

   b. What is the molarity of a solution of AZ if its K_{sp} is the same value?
   ans: $[AZ] = 6.52 \times 10^{-4}$ M

Sample Problem 3

Will precipitation of strontium sulfate occur when 50.0 mL of 0.025 M strontium nitrate solution is mixed with 50.0 mL of a 0.014 M copper(II) sulfate solution? The K_{sp} of strontium nitrate is 3.2×10^{-7}.

Solution

ANALYZE

What is given in the problem? **the molar concentrations of the solutions to be mixed, the identities of the solutes, and the volumes of the solutions to be mixed**

What are you asked to find? **whether a precipitate of strontium sulfate forms when the two solutions are mixed**

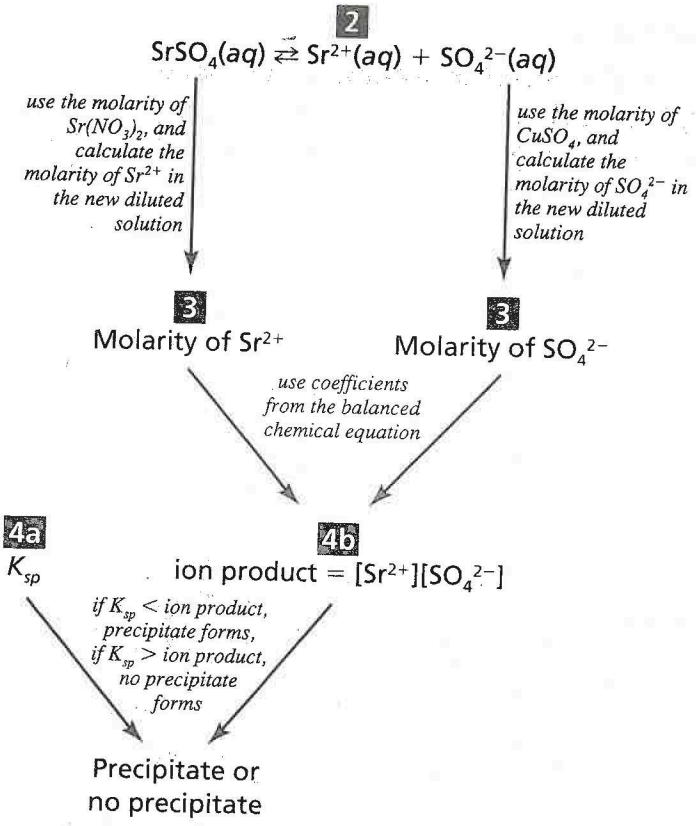
Items	Data
Concentration of solution 1	0.025 M $SrNO_3$
Volume of solution 1	50.0 mL
Concentration of solution 2	0.014 M $CuSO_4$
Volume of solution 2	50.0 mL
Volume of combined solution	100.0 mL
Concentration of combined solution	? M $SrSO_4$
Potential precipitate	$SrSO_4$
K_{sp} of $SrSO_4$	3.2×10^{-7}
Precipitate/no precipitate forms	?

PLAN

What steps are needed to determine whether a precipitate will form?

Calculate the molar concentrations of the ions that can form a precipitate in the new volume of solution. Use these concentrations to calculate the ion product. Compare the ion product with K_{sp}.

Problem Solving *continued*



Write the balanced equation for the dissociation of $SrSO_4$.

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

Calculate the molarities of Sr^{2+} and SO_4^{2-}. This is a simple dilution calculation. The subscript 1 in each case represents that individual solution; the subscript 2 represents the combined solution.

$$Sr(NO_3)_2(aq) \rightleftarrows Sr^{2+}(aq) + 2NO_3^-(aq)$$
$$[Sr^{2+}]_1 = [Sr(NO_3)_2]$$
$$[Sr^{2+}]_1 V_1 = [Sr^{2+}]_2 V_2$$

$$[Sr^{2+}]_2 = \frac{\overset{\text{calculated above}}{[Sr^{2+}]_1} \overset{\text{given}}{V_1}}{\underset{\text{sum of volumes of solutions mixed}}{V_2}}$$

$$CuSO_4(aq) \rightleftarrows Cu^{2+}(aq) + SO_4^{2-}(aq)$$
$$[SO_4^{2-}]_1 = [CuSO_4]$$
$$[SO_4^{2-}]_1 V_1 = [SO_4^{2-}]_2 V_2$$

$$[SO_4^{2-}]_2 = \frac{\overset{\text{calculated above}}{[SO_4^{2-}]_1} \overset{\text{given}}{V_1}}{\underset{\text{sum of volumes of solutions mixed}}{V_2}}$$

Calculate the ion product for $SrSO_4$.

$$\text{ion product} = [Sr^{2+}]_{\substack{\text{calculated} \\ \text{above}}}[SO_4^{2-}]_2$$

Compare the ion product to the K_{sp} value to determine if precipitation occurs.

COMPUTE

$$[Sr^{2+}]_2 = \frac{0.025\ \text{M} \times 50.0\ \cancel{\text{mL}}}{100.0\ \cancel{\text{mL}}} = 1.2 \times 10^{-2}\ \text{M}$$

$$[SO_4^{2-}]_2 = \frac{0.014\ \text{M} \times 50.0\ \cancel{\text{mL}}}{100.0\ \cancel{\text{mL}}} = 7.0 \times 10^{-3}\ \text{M}$$

$$\text{ion product} = [1.2 \times 10^{-2}][7.0 \times 10^{-3}] = 8.4 \times 10^{-5}$$

$$K_{sp} = 3.2 \times 10^{-7}$$

$$K_{sp} < \text{ion product}$$

Precipitation will occur.

EVALUATE

Are the units correct?

Yes; the ion product has no units.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because data were given to a minimum of two significant figures.

Is the answer reasonable?

Yes; the calculation can be approximated as $0.01 \times 0.007 = 0.000\ 07 = 7 \times 10^{-5}$, which is of the same order of magnitude as the calculated result.

Practice

In each of the following problems, include the calculated ion product with your answer.

1. Will a precipitate of $Ca(OH)_2$ form when 320. mL of a 0.046 M solution of NaOH mixes with 400. mL of a 0.085 M $CaCl_2$ solution? K_{sp} of $Ca(OH)_2$ is 5.5×10^{-6}. **ans: ion product = 1.9×10^{-5}, precipitation occurs.**

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Problem Solving *continued*

2. 20.00 mL of a 0.077 M solution of silver nitrate, $AgNO_3$, is mixed with 30.00 mL of a 0.043 M solution of sodium acetate, $NaC_2H_3O_2$. Does a precipitate form? The K_{sp} of $AgC_2H_3O_2$ is 2.5×10^{-3}.
ans: ion product = 8.1×10^{-4}, no precipitate

3. If you mix 100. mL of 0.036 M $Pb(C_2H_3O_2)_2$ with 50. mL of 0.074 M NaCl, will a precipitate of $PbCl_2$ form? K_{sp} of $PbCl_2$ is 1.9×10^{-4}.
ans: ion product = 1.5×10^{-5}, no precipitate

4. If 20.00 mL of a 0.0090 M solution of $(NH_4)_2S$ is mixed with 120.00 mL of a 0.0082 M solution of $Al(NO_3)_3$, does a precipitate form? The K_{sp} of Al_2S_3 is 2.00×10^{-7}. **ans: ion product = 1.1×10^{-13}, no precipitate**

Additional Problems—Equilibrium of Salts, K_{sp}

1. The molar concentration of a saturated calcium chromate, $CaCrO_4$, solution is 0.010 M at 25°C. What is the K_{sp} of calcium chromate?
2. A 10.00 mL sample of a saturated lead selenate solution is found to contain 0.00136 g of dissolved $PbSeO_4$ at 25°C. Determine the K_{sp} of lead selenate.
3. A 22.50 mL sample of a saturated copper(I) thiocyanate, CuSCN, solution at 25°C is found to have a 4.0×10^{-6} M concentration.
   a. Determine the K_{sp} of CuSCN.
   b. What mass of CuSCN would be dissolved in 1.0×10^{3} L of solution?
4. A saturated solution of silver dichromate, $Ag_2Cr_2O_7$, has a concentration of 3.684×10^{-3} M. Calculate the K_{sp} of silver dichromate.
5. The K_{sp} of barium sulfite, $BaSO_3$, at 25°C is 8.0×10^{-7}.
   a. What is the molar concentration of a saturated solution of $BaSO_3$?
   b. What mass of $BaSO_3$ would dissolve in 500. mL of water?
6. The K_{sp} of lead(II) chloride at 25°C is 1.9×10^{-4}. What is the molar concentration of a saturated solution at 25°C?
7. The K_{sp} of barium carbonate at 25°C is 1.2×10^{-8}.
   a. What is the molar concentration of a saturated solution of $BaCO_3$ at 25°C?
   b. What volume of water would be needed to dissolve 0.10 g of barium carbonate?
8. The K_{sp} of $SrSO_4$ is 3.2×10^{-7} at 25°C.
   a. What is the molar concentration of a saturated $SrSO_4$ solution?
   b. If 20.0 L of a saturated solution of $SrSO_4$ were evaporated to dryness, what mass of $SrSO_4$ would remain?
9. The K_{sp} of strontium sulfite, $SrSO_3$, is 4.0×10^{-8} at 25°C. If 1.0000 g of $SrSO_3$ is stirred in 5.0 L of water until the solution is saturated and then filtered, what mass of $SrSO_3$ would remain?
10. The K_{sp} of manganese(II) arsenate is 1.9×10^{-11} at 25°C. What is the molar concentration of $Mn_3(AsO_4)_2$ in a saturated solution? Note that five ions are produced from the dissociation of $Mn_3(AsO_4)_2$.
11. Suppose that 30.0 mL of a 0.0050 M solution of $Sr(NO_3)_2$ is mixed with 20.0 mL of a 0.010 M solution of K_2SO_4 at 25°C. The K_{sp} of $SrSO_4$ is 3.2×10^{-7}.
    a. What is the ion product of the ions that can potentially form a precipitate?
    b. Does a precipitate form?
12. Lead(II) bromide, $PbBr_2$, is slightly soluble in water. Its K_{sp} is 6.3×10^{-6} at 25°C. Suppose that 120. mL of a 0.0035 M solution of $MgBr_2$ is mixed with 180. mL of a 0.0024 M $Pb(C_2H_3O_2)_2$ solution at 25°C.
    a. What is the ion product of Br^- and Pb^{2+} in the mixed solution?
    b. Does a precipitate form?

13. The K_{sp} of $Mg(OH)_2$ at 25°C is 1.5×10^{-11}.

a. Write the equilibrium equation for the dissociation of $Mg(OH)_2$.

b. What volume of water would be required to dissolve 0.10 g of $Mg(OH)_2$?

c. Considering that magnesium hydroxide is essentially insoluble, why is it possible to titrate a suspension of $Mg(OH)_2$ to an equivalence point with a strong acid such as HCl?

14. Lithium carbonate is somewhat soluble in water; its K_{sp} at 25°C is 2.51×10^{-2}.

a. What is the molar concentration of a saturated Li_2CO_3 solution?

b. What mass of Li_2CO_3 would you dissolve in order to make 3440 mL of saturated solution?

15. A 50.00 mL sample of a saturated solution of barium hydroxide, $Ba(OH)_2$, is titrated to the equivalence point by 31.61 mL of a 0.3417 M solution of HCl. Determine the K_{sp} of $Ba(OH)_2$.

16. Calculate the K_{sp} for salts represented by QR that dissociate into two ions, Q^+ and R^-, in each of the following solutions:

a. saturated solution of QR is 1.0 M

b. saturated solution of QR is 0.50 M

c. saturated solution of QR is 0.1 M

d. saturated solution of QR is 0.001 M

17. Suppose that salts QR, X_2Y, KL_2, A_3Z, and D_2E_3 form saturated solutions that are 0.02 M in concentration. Calculate K_{sp} for each of these salts.

18. The K_{sp} at 25°C of silver bromide is 5.0×10^{-13}. What is the molar concentration of a saturated AgBr solution? What mass of silver bromide would dissolve in 10.0 L of saturated solution at 25°C?

19. The K_{sp} at 25°C for calcium hydroxide is 5.5×10^{-6}.

a. Calculate the molarity of a saturated $Ca(OH)_2$ solution.

b. What is the OH^- concentration of this solution?

20. The K_{sp} of magnesium carbonate is 3.5×10^{-8} at 25°C. What mass of $MgCO_3$ would dissolve in 4.00 L of water at 25°C?