

Chapter 18

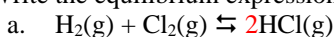
Practice Test

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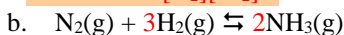
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1. Reactions that may proceed in a forward or reverse reaction are called reversible reactions. They eventually reach a point where the forward and reverse reactions occur at the same rate. This point is called equilibrium. Once the reactions reach this point, the concentrations of reactants and products become constant, but rarely equal.
2. An equilibrium expression is a ratio of products to reactants at equilibrium. Its value is represented by the symbol K_{eq} . The ratio is expressed by placing the concentrations of products of the reaction in the numerator and the concentrations of reactants in the denominator. The coefficients from the balanced equation are used as exponents. Because their concentrations never change at a given temperature, solids and liquids are not included in the equilibrium expression.

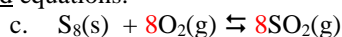
3. Write the equilibrium expression for each of the following unbalanced equations.



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

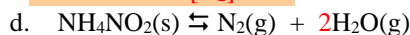


$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$



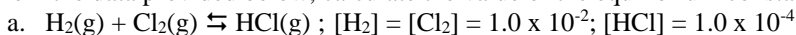
$$K_{eq} = \frac{[SO_2]^8}{[O_2]^8}$$

[S₈] doesn't appear because it's a solid



$$K_{eq} = [N_2][H_2O]^2$$

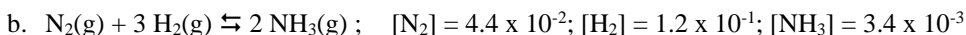
4. From the data provided below, calculate the value of the equilibrium constant for the reaction.



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

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

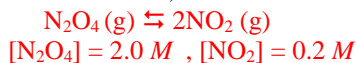
[NH₄NO₂] doesn't appear because it's a solid



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

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

5. To check if reaction has reached equilibrium a reaction quotient, represented by the symbol Q, may be calculated. If $Q < K_{eq}$, then the reaction will continue to proceed forward, if $Q > K_{eq}$, the reaction will proceed in reverse, and if $Q = K_{eq}$, the reaction is at equilibrium.
6. For the reaction $N_2O_4(g) \rightleftharpoons 2NO_2(g)$, $K_{eq} = 0.2$. 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?

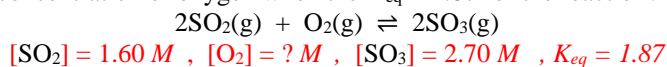


$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$

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

$K_{eq} = 0.2$ $Q < K_{eq}$, so the reaction will proceed to the right (forward).

7. For the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, the equilibrium concentrations of the sulfur oxides are $[SO_2] = 1.60$ and $[SO_3] = 2.70$. What is the concentration of oxygen when the $K_{eq} = 1.87$ for the reaction?



$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$O_2 = \frac{[SO_3]^2}{[SO_2]^2 K_{eq}}$$

$$O_2 = \frac{[2.70]^2}{[1.60]^2 \cdot 1.87} = 1.52 M$$

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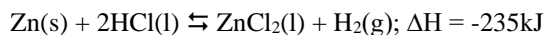
8. Name the 3 stresses that may be applied to a chemical equilibrium to cause it to shift.

_____ temperature _____ pressure or volume _____ concentration _____

9. 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.

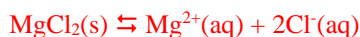
10. For the reaction below, mark whether the stress listed will cause the reaction to move forward or in reverse.



- | | | | | | |
|-------------------------------|-------|----------------------------------|-------|-------------------------------|-------|
| 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. | c. Reverse to consume added H ₂ . | i. Forward to replace H ₂ . |
| b. Reverse reaction requires less volume (no gases produced), so reaction driven in reverse. | d. Forward to consume Zn. | j. Forward. |
| | e. Reverse to consume ZnCl ₂ . | k. Forward. |
| | f. Forward to consume HCl. | |
| | g. Reverse to replace HCl. | |
| | h. Forward to replace ZnCl ₂ . | |

11. 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 MgCl₂ in water, this is what happens:



Ions of both Mg²⁺ and Cl⁻ would be floating around in the water. If you then add a second solution that has NaCl [NaCl(s) ⇌ Na⁺(aq) + Cl⁻(aq)], the added Cl⁻ would cause the above reaction to shift in reverse and MgCl₂(s) would precipitate out of the solution. The "common ion" affects the equilibrium of one of the reactions (LeChatelier's Principle).

12. Solution formation also reaches a dynamic equilibrium. The formation of a solution can be described using a dissociation equation. In these equations, the reactant is always a solid and the products are always aqueous ions. The solubility of a substance can be described using a solubility product expression, which is just an equilibrium expression for a dissolution reaction. In these expressions there will never be a denominator. Like K_{eq}, the value of K_{sp} is temperature dependent.

13. Write the balanced dissociation equations for the following salts.

- | | | | |
|--|--|--------------------------------------|--|
| a. CaCl ₂ | <u>CaCl₂(s) ⇌ Ca²⁺(aq) + 2Cl⁻(aq)</u> | c. NaI | <u>NaI(s) ⇌ Na⁺(aq) + I⁻(aq)</u> |
| b. (NH ₄) ₂ SO ₄ | <u>(NH₄)₂SO₄(s) ⇌ 2NH₄⁺(aq) + SO₄²⁻(aq)</u> | d. Al(NO ₃) ₃ | <u>Al(NO₃)₃(s) ⇌ Al³⁺(aq) + 3NO₃⁻(aq)</u> |

14. Q_{sp} is called the solubility product. If Q_{sp} > K_{sp}, the solution is supersaturated. If Q_{sp} < K_{sp}, the solution is unsaturated. If Q_{sp} = K_{sp}, the solution is saturated.

15. Write the dissociation equations and solubility product expressions for the following:

- | | | | |
|----------------------|--|--|--|
| a. SrSO ₄ | <u>SrSO₄(s) ⇌ Sr²⁺(aq) + SO₄²⁻(aq)</u>
<u>K_{sp} = [Sr²⁺][SO₄²⁻]</u> | b. Al ₂ (SO ₄) ₃ | <u>Al₂(SO₄)₃(s) ⇌ 2Al³⁺(aq) + 3SO₄²⁻(aq)</u>
<u>K_{sp} = [Al³⁺]²[SO₄²⁻]³</u> |
|----------------------|--|--|--|

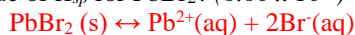
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16. A sample of $\text{PbBr}_2(\text{s})$ is added to pure water and allowed to come to equilibrium at 25°C . The concentration of Pb^{2+} is 0.0118M at equilibrium. What is the value of K_{sp} for PbBr_2 ? (6.60×10^{-6})



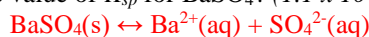
$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{Br}^-]^2$$

The concentrations of Br^- will be 2x the concentration of Pb^{2+} because the mole ratio of Br^- to Pb^{2+} is 2:1

$$K_{\text{sp}} = [0.0118][0.0236]^2$$

$$K_{\text{sp}} = 6.6 \times 10^{-6}$$

17. A sample of $\text{BaSO}_4(\text{s})$ is added to pure water and allowed to come to equilibrium at 25°C . The concentration of Ba^{2+} is $1.05 \times 10^{-5}\text{M}$ at equilibrium. What is the value of K_{sp} for BaSO_4 ? (1.1×10^{-10})



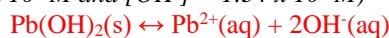
$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

The concentrations of Ba^{2+} and SO_4^{2-} will be equal because their mole ratio is 1:1

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

$$K_{\text{sp}} = 1.10 \times 10^{-10}$$

18. What will be the equilibrium concentration of dissolved ions in a saturated solution of $\text{Pb}(\text{OH})_2$ at 25°C ? K_{sp} for the reaction is 1.2×10^{-15} . ($[\text{Pb}^{2+}] = 6.69 \times 10^{-6}\text{M}$ and $[\text{OH}^-] = 1.34 \times 10^{-5}\text{M}$)



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

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:

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

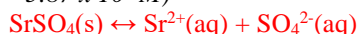
$$1.2 \times 10^{-15} = 4x^3$$

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

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

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

19. 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} . ($[\text{Sr}^{2+}] = [\text{SO}_4^{2-}] = 5.87 \times 10^{-4}\text{M}$)



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

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, $[\text{Sr}^{2+}]$ is $5.87 \times 10^{-4}\text{M}$ and $[\text{SO}_4^{2-}]$ would be the same.