

AP Chemistry Unit 6 Review

Equilibrium

Honors Lecture Notes: Unit 5

Definitions

Most reactions are not one way; typically both the forward and the reverse reactions are occurring simultaneously. When the rates of the forward and reverse reaction rates are equal, equilibrium is achieved. This does not mean the reaction stops, but it does mean the concentration of the various species is remaining constant.

Equilibrium Constant

The equilibrium constant, K_{eq} , is calculated as follows:

For: $A(s) + 2 B(aq) \rightleftharpoons C(g) + 3 D(aq)$ $K_{eq} = \frac{[C][D]^3}{[B]^2}$

Note that solids and liquids do not appear in K_{eq} . If $K_{eq} > 1$, the reaction is favored. If $K_{eq} < 1$, the reaction is not favored. There are many variants on the equilibrium constant. The two most common are K_c , which utilized concentrations, and K_p , which utilizes partial pressures.

Manipulating K_{eq}

When an equation is manipulated, the K_{eq} values also changes. There are three rules regarding this:

- 1) If you flip the reaction, you take the inverse of K_{eq}
- 2) If you multiply the coefficients on the reaction by any value, you turn that value into an exponent of the K_{eq}
- 3) If you combine two reactions, the K_{eq} of the new reaction will be the product for the K_{eq} values of the initial reactions

Given: $H_2(g) + S(s) \rightleftharpoons H_2S(g)$ $K_{eq} = 9.8 \times 10^5$

$H_2S(g) \rightleftharpoons H_2(g) + S(s)$ $K_{eq} = 1/(9.8 \times 10^5) = 1.0 \times 10^{-6}$

$2 H_2(g) + 2 S(s) \rightleftharpoons 2 H_2S(g)$ $K_{eq} (9.8 \times 10^5)^2 = 9.8 \times 10^{10}$

Gibbs and K_{eq}

The equation which relates Gibbs free energy and K_{eq} is:

$$\Delta G = -RT \ln K$$

ΔG is in J/mol

R is 8.31 J/mol * K

T is in Kelvins

K is unitless

As a summary, the three ways favorability (spontaneity) is measured are:

	Favored	Non-Favored
ΔG	Negative	Positive
K	Greater than 1	Less than 1
E°_{cell}	Positive	Negative

Le Chatelier's Principle

When a system at equilibrium is exposed to stress, the system will "shift" in response to that stress in order to re-establish equilibrium. A shift is the increase of either the forward (right shift) or reverse (left shift) reaction rate. There are multiple types of stressors that can cause equilibrium shifts.

- 1) Concentration. If you increase the concentration of a product or reactant, the reaction will shift away from the increase. If you decrease the concentration of a product or reactant, the reaction will shift towards the decrease. Remember, solids and liquids do not appear in equilibrium expressions, so changing the amount of either of these will not change the concentrations or cause shifts.
- 2) Temperature. For exothermic reactions, heat is a product, and for endothermic reactions, heat is a reactant. Increasing temp shifts a reaction away from the side with the heat, and lowering temp shifts a reaction towards the side with the heat.
- 3) Pressure. If the pressure on a reaction is increased (often by decreasing the volume of the container), it shifts to the side with less gas molecules. If the pressure on a reaction is decrease, it shifts to the side with more gas molecules.

The only type of stressor that can actually change the value of the equilibrium constant is a temperature change, and the value would change in accordance with the shift (right shift = increased K and left shift = decreased K). Concentration and pressure stressors do not change the value for K , and the equilibrium concentrations will eventually be reestablished.

Review Sheet Unit 6 Chemistry Answer Key

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Table of Contents Review Sheet Unit 6 Chemistry Answer Key

1. Understanding the eBook Review Sheet Unit 6 Chemistry Answer Key
 - The Rise of Digital Reading Review Sheet Unit 6 Chemistry Answer Key
 - Advantages of eBooks Over Traditional Books
2. Identifying Review Sheet Unit 6 Chemistry Answer Key
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Review Sheet Unit 6 Chemistry Answer Key
 - User-Friendly Interface
4. Exploring eBook Recommendations from Review Sheet Unit 6 Chemistry Answer Key
 - Personalized Recommendations
 - Review Sheet Unit 6 Chemistry Answer Key User Reviews and Ratings
 - Review Sheet Unit 6 Chemistry Answer Key and Bestseller Lists
5. Accessing Review Sheet Unit 6 Chemistry Answer Key Free and Paid eBooks
 - Review Sheet Unit 6 Chemistry Answer Key Public Domain eBooks
 - Review Sheet Unit 6 Chemistry Answer Key eBook Subscription Services
 - Review Sheet Unit 6 Chemistry Answer Key Budget-Friendly Options
6. Navigating Review Sheet Unit 6 Chemistry Answer Key eBook Formats

- ePub, PDF, MOBI, and More
 - Review Sheet Unit 6 Chemistry Answer Key Compatibility with Devices
 - Review Sheet Unit 6 Chemistry Answer Key Enhanced eBook Features
7. Enhancing Your Reading Experience
- Adjustable Fonts and Text Sizes of Review Sheet Unit 6 Chemistry Answer Key
 - Highlighting and Note-Taking Review Sheet Unit 6 Chemistry Answer Key
 - Interactive Elements Review Sheet Unit 6 Chemistry Answer Key
8. Staying Engaged with Review Sheet Unit 6 Chemistry Answer Key
- Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Review Sheet Unit 6 Chemistry Answer Key
9. Balancing eBooks and Physical Books Review Sheet Unit 6 Chemistry Answer Key
- Benefits of a Digital Library
 - Creating a Diverse Reading Collection Review Sheet Unit 6 Chemistry Answer Key
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Review Sheet Unit 6 Chemistry Answer Key
- Setting Reading Goals Review Sheet Unit 6 Chemistry Answer Key
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Review Sheet Unit 6 Chemistry Answer Key
- Fact-Checking eBook Content of Review Sheet Unit 6 Chemistry Answer Key
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
- Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
- Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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