

## Unit 5 – Gases and Gas Laws

### I. The Nature of Gases

#### I. The Kinetic Theory and a Molecular Model of Gases

##### A. Assumptions of the Kinetic Molecular Theory

1. Gases consist of large numbers of tiny particles that are far apart relative to their size.
2. The particles undergo a lot of collisions.
  - a. Collisions in which no energy is lost.
3. The particles are in constant, random motion. They therefore possess kinetic energy, the energy of motion.

#### II. Gas Pressure

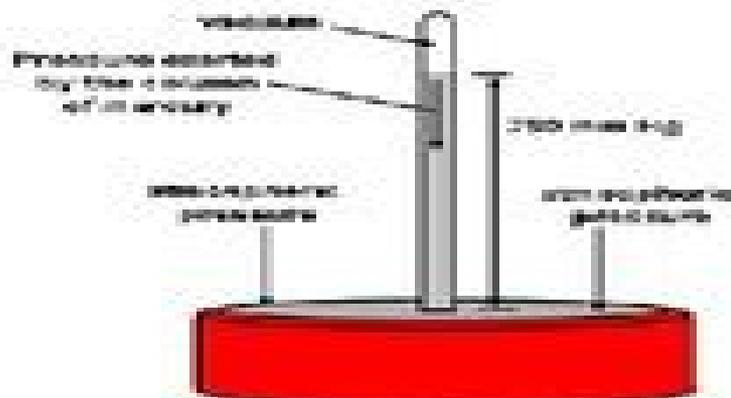
##### A. Pressure

1. The force per unit area on a surface

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

2. The particles exert force, and therefore pressure, on any surface with which they collide.

##### B. Unit of Pressure



**Unit of Pressure**

Unit	Symbol	Equivalent Relationship
Force/area	N/m <sup>2</sup>	SI pressure unit
		1 Pa = 1 N/m <sup>2</sup>
Millimeter of mercury	mm Hg	Pressure that supports a 1 mm column of mercury in a barometer
Atmosphere	atm	Pressure that supports a column of mercury about 760 mm high
Torr	torr	1 torr = 1 mm Hg

##### C. Standard Pressure

1. 1 atm = 760 mm Hg = 760 torr = 101.3 kPa

### III. Kinetic Energy and Temperature

#### A. Formula for Kinetic Energy

$$KE = \frac{1}{2}mv^2 \quad m = \text{mass} \quad v = \text{speed}$$

#### B. Relationship to Temperature

1. The average kinetic energy of gas particles depends on the temperature.
2. All gases at the same temperature have the same average kinetic energy.
  - a. Heavier molecules (mass  $m$ ) have higher average speeds.
3. Kinetic temperature is directly proportional to the average kinetic energy of a molecule.
  - a. 0 K = 0 atm = absolute zero = 0 K kinetic energy

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## **Table of Contents Holt Chemistry Questions For Section 5 1**

1. Understanding the eBook Holt Chemistry Questions For Section 5 1
  - The Rise of Digital Reading Holt Chemistry Questions For Section 5 1
  - Advantages of eBooks Over Traditional Books
2. Identifying Holt Chemistry Questions For Section 5 1
  - 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 Holt Chemistry Questions For Section 5 1
  - User-Friendly Interface
4. Exploring eBook Recommendations from Holt Chemistry Questions For Section 5 1
  - Personalized Recommendations
  - Holt Chemistry Questions For Section 5 1 User Reviews and Ratings
  - Holt Chemistry Questions For Section 5 1 and Bestseller Lists
5. Accessing Holt Chemistry Questions For Section 5 1 Free and Paid eBooks
  - Holt Chemistry Questions For Section 5 1 Public Domain eBooks
  - Holt Chemistry Questions For Section 5 1 eBook Subscription Services

- Holt Chemistry Questions For Section 5 1 Budget-Friendly Options
- 6. Navigating Holt Chemistry Questions For Section 5 1 eBook Formats
  - ePub, PDF, MOBI, and More
  - Holt Chemistry Questions For Section 5 1 Compatibility with Devices
  - Holt Chemistry Questions For Section 5 1 Enhanced eBook Features
- 7. Enhancing Your Reading Experience
  - Adjustable Fonts and Text Sizes of Holt Chemistry Questions For Section 5 1
  - Highlighting and Note-Taking Holt Chemistry Questions For Section 5 1
  - Interactive Elements Holt Chemistry Questions For Section 5 1
- 8. Staying Engaged with Holt Chemistry Questions For Section 5 1
  - Joining Online Reading Communities
  - Participating in Virtual Book Clubs
  - Following Authors and Publishers Holt Chemistry Questions For Section 5 1
- 9. Balancing eBooks and Physical Books Holt Chemistry Questions For Section 5 1
  - Benefits of a Digital Library
  - Creating a Diverse Reading Collection Holt Chemistry Questions For Section 5 1
- 10. Overcoming Reading Challenges
  - Dealing with Digital Eye Strain
  - Minimizing Distractions
  - Managing Screen Time
- 11. Cultivating a Reading Routine Holt Chemistry Questions For Section 5 1
  - Setting Reading Goals Holt Chemistry Questions For Section 5 1
  - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Holt Chemistry Questions For Section 5 1
  - Fact-Checking eBook Content of Holt Chemistry Questions For Section 5 1
  - 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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