Chapter 5 Reading Guide **Name:**
AP Chemistry 2016-2017 **Date:** **Per:**

Gases and their properties are the focus of this chapter. These concepts are in Big Ideas 2, 3, and 4 of the AP Chemistry curriculum. Also included in this chapter are cross domain applications of gases, the focus of science practice 7.

5.1 Breathing: Putting Pressure to Work

1. Define pressure.

2. How do gases apply pressure on objects?

3. How does the concentration of a gas affect pressure?

5.2 Pressure: The Result of Molecular Collisions

4. Explain why the pressure of a gas is directly proportional to the number of gas molecules.

5. Explain why pressure can be measured in millimeters mercury, even though millimeters measure length.

6. Draw two diagrams representing small volumes or cross-sections of nitrogen gas, at the particulate level, at 760 mmHg and 700 mmHg.

7. Climbers on Mount Everest often die from altitude sickness, caused by low gas pressure. Why is air pressure less on the summit of Everest, compared to the air pressure at sea level?

8. What are some common units used to measure air pressure?

5.3 The Simple Gas Laws: Boyles's Law, Charles's Law, and Avogadro's Law

9. What are the four properties of a gas that can be measured?

10. Fill out the table below for the three simple gas laws.

Gas Law	Variables in equation	Constant variables	Equation	Sketch graph
Boyle's				
Charles's				
Avogadro's				

11. Give a real-world example illustrating how Boyle's Law works.

12. What is the connection between absolute zero and Charles's Law?

13. Give a real-world example showing how Charles's Law works.

14. When solving Charles's Law problems, what temperature units must be used?

15. Give an example of Avogadro's Law.

16. Why do warm gas molecules rise?

5.4 The Ideal Gas Law

17. Which simple gas laws combine to make the Ideal Gas Law?

18. Explain what R is, write the numerical value, and include the units when working with gases.

19. Give the correct units for each variable in the ideal gas law:
    a. P

    b. V

    c. n

    d. T

20. What two properties are variables in Gay-Lussac's Law? Which properties must be constant?

21. Explain why you should not heat up an aerosol can.

5.5 Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas

22. Define "molar volume".

23. To make comparisons among gas samples, the properties of gases are often measured or calculated at STP. What is STP, and what are its conditions?

24. Is there an equivalent to STP for comparing liquids or solids? Explain why or why not.

25. At STP, what is the molar volume of an ideal gas?

26. How is the density of a gas calculated at STP using its molar volume?

27. What gases in a balloon would cause the balloon to rise in air? What gases would cause the balloon to sink? Explain.

28. Combine the equation for molar mass and ideal gases to write one equation for calculating the molar mass of a gas.

5.6 Mixtures of Gases and Partial Pressures

29. Define the term “partial pressure of a gas”.

30. How is the partial pressure of a gas determined?

31. In a mixture of gases, if the partial pressure of each gas is known, what is the equation used to determine the total pressure? What is this relationship called?

32. What does “mole fraction” mean?

33. What equation is used to calculated the mole fraction of a gas?

34. In the chemistry laboratory how are gases from chemical reactions often collected?

35. When collecting a gas through water displacement, what needs to be done to determine the pressure of the gas being collected? What other value needs to be measured?

36. In a chemistry lab, what common reaction is used to produce hydrogen gas? Why is the gas collected “over water”?

37. Write an equation to use for the calculation used to answer question #36.

38. How is the vapor pressure of water determined?

39. **Explain** (not describe) the difference in the vapor pressure of water at 25 °C and 55 °C.

5.7 Gases in Chemical Reactions: Stoichiometry Revisited

40. How are stoichiometry problems with gases different from other stoichiometry problems?

41. What conditions must be met in order to use 22.4 L/mol as the molar volume of a gas?

5.8 Kinetic Molecular Theory: A Model for Gases

42. Summarize the three assumptions of kinetic molecular theory.

43. Why is it usually safe to assume that gas particles size does not influence the volume of a gas?

44. If two gases are at the same temperature, what can we assume about them?

45. Differentiate between elastic and inelastic collisions of gas particles.

46. Explain each of the following gas laws using KMT
    a. Boyle's Law

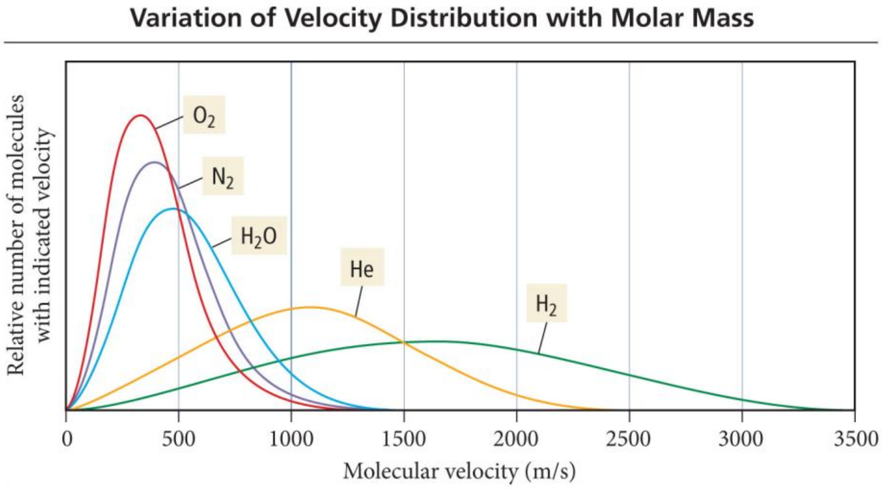
    b. Charles's Law

    c. Avogadro's Law

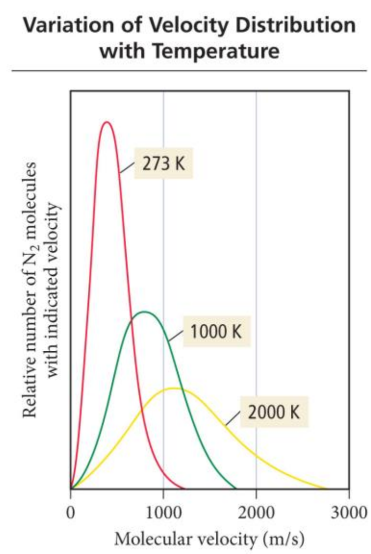
    d. Dalton's Law

47. In a mixture of gases at the same temperature, how are the masses of gas particles related to their velocities?

48. Summarize the key concepts shown in figure 5.18, shown below.



49. Describe two key concepts shown in figure 5.19, shown at right.



5.9 Mean Free Path, Diffusion, and Effusion of Gases

50. Explain the concept of "mean-free path" for a gas.

51. What is diffusion of a gas, and how does molar mass of a gas affect diffusion?

52. Define the process of effusion, and explain how molar mass influences the rate of effusion.

53. Explain Graham's Law using concepts of kinetic molecular theory.

54. Two balloons are inflated to the same volume, one with air (mostly N_2, nitrogen) and one with hydrogen, H_2. What differences in volume over time would you observe between the two balloons? Explain why.

5.10 Real Gases: The Effects of Size and Intermolecular Forces (*note that you are NOT responsible for the Van der Waals equation. Focus on key concepts in this section*)

55. What are two differences between a real gas and an ideal gas?

56. Under what two conditions do the ideal gas assumptions break down?

57. At high pressure, how will the volume of a real gas change? Explain why.

58. How do real gas molecular attractions affect the pressure of gases? Explain why.