CHAPTER 11 REVIEW

CHAPTER SUMMARY

11.1* Gay-Lussac’s law of combining volumes of gases states that the volumes of reacting gases and their products at the same temperature and pressure can be expressed as ratios of small whole numbers.
• Gay-Lussac’s law and Avogadro’s law can be used to show that molecules of the reactive elemental gases are diatomic.
• The volume occupied by one mole of an ideal gas at STP is called the standard molar volume. The standard molar volume of an ideal gas is 22.414 L at standard temperature and pressure.

Vocabulary
Avogadro’s law (334) Gay-Lussac’s law of combining volumes of gases (333) standard molar volume of a gas (333)

11.2* Charles’s law, Boyle’s law, and Avogadro’s law can be combined to create the ideal gas law. The ideal gas law is stated mathematically as follows: $PV = nRT$.
• The value and units of the ideal gas constant depend on the units of the variables used in the ideal gas law.

Vocabulary
ideal gas constant (342) ideal gas law (340)

11.3* Given a balanced equation and the volume of a reacting gas, the volume of another gaseous product or reactant can be calculated using their volume ratios, as long as reactants and products exist under the same conditions.
• Given the volume of a reactant or product gas, the mass of a second reactant or product can be calculated using the ideal gas law and molar-to-mass conversion factors.
• If the mass of a substance is known, the ideal gas law and the proper mass-to-mole conversion factors can be used to calculate the volume of a gas.

Vocabulary
Graham’s law of effusion (352)

11.4* Graham’s law of effusion states that the relative rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.
• Graham’s law reflects the fact that less massive molecules effuse faster than do more massive ones.

Vocabulary
Graham’s law of effusion (352)

REVIEWING CONCEPTS

1. a. What restrictions are there on the use of Gay-Lussac’s law of combining volumes?
b. At the same temperature and pressure, what is the relationship between the volume of a gas and the number of molecules present? (11-1)

2. According to Avogadro,
• a. what is the relationship between gas volume and number of moles at constant temperature and pressure?
b. what is the mathematical expression denoting this relationship? (11-1)

3. What is the relationship between the number of molecules and the mass of 22.4 L of different gases at STP? (11-1)

4. Why must the temperature and pressure be specified when stating gas density values? (11-1)

5. a. Write the equation for the ideal gas law.
b. What relationship is expressed in the ideal gas law? (11-2)

6. a. In what situation does the ideal gas law apply?
b. Why do you have to pay particular attention to units when using this law? (11-2)

7. a. In a balanced chemical equation, what is the relationship between the molar ratios and the volume ratios of gaseous reactants and products?
b. What restriction applies to the use of the volume ratios in solving stoichiometry problems? (11-3)

8. a. Distinguish between diffusion and effusion.
b. At a given temperature, what factor determines the rates at which different molecules undergo these processes? (11-4)

PROBLEMS

Molar Volume and Gas Density

9. Suppose a 5.00 L sample of O2 at a given temperature and pressure contains $1.08 \times 10^{23}$ molecules. How many molecules would be contained in each of the following at the same temperature and pressure?
• a. 5.00 L H2

10. How many molecules are contained in each of the following?
• a. 1.00 mol O2
• b. 2.50 mol He
c. 0.0050 mol NH3
d. 11.5 g NO2

11. Find the mass of each of the following.
• a. 2.25 mol CO2
• b. 3.01 $\times 10^{22}$ molecules H2S
c. 25.0 molecules SO2

12. What is the volume, in liters, of each of the following at STP? (Hint: See Sample Problem 11-1.)
• a. 1.00 mol O2
• b. 3.50 mol F2
c. 0.0400 mol CO2
d. 1.20 $\times 10^4$ mol He

13. How many moles are contained in each of the following at STP?
• a. 22.4 L N2
• b. 5.60 L Cl2
c. 0.125 L Ne
d. 70.0 mL NH3

14. Find the mass, in grams, of each of the following at STP. (Hint: See Sample Problem 11-2.)
• a. 11.2 L H2
• b. 2.80 L CO2
• c. 15.0 mL SO2
d. 3.40 cm3 F2

15. Find the volume, in liters, of each of the following at STP.
• a. 8.00 g O3
• b. 3.50 g CO
c. 0.0760 g H2S
d. 2.25 $\times 10^3$ kg NH3

Ideal Gas Law

16. Calculate the pressure, in atmospheres, exerted by each of the following. (Hint: See Sample Problem 11-3.)
• a. 2.50 L of HF containing 1.35 mol at 320 K
• b. 4.75 L of NO2 containing 0.86 mol at 300 K
c. 7.50 $\times 10^3$ mL of CO2 containing 2.15 mol at 57°C

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17. Calculate the volume, in liters, occupied by each of the following. (Hint: See Sample Problem 11-4.)
   a. 2.00 mol of H₂ at 300 K and 1.25 atm
   b. 0.425 mol of NH₃ at 37°C and 0.724 atm
   c. 4.00 g of O₂ at 57°C and 0.888 atm

18. Determine the number of moles of gas contained in each of the following:
   a. 1.25 L at 250 K and 1.06 atm
   b. 0.80 L at 27°C and 0.925 atm
   c. 7.50 x 10⁻³ mL at –50°C and 0.921 atm

19. Find the mass of each of the following. (Hint: See Sample Problem 11-5.)
   a. 5.60 L of O₂ at 1.75 atm and 250 K
   b. 3.50 L of NH₃ at 0.921 atm and 27°C
   c. 125 mL of SO₂ at 0.822 atm and –53°C

20. Find the molar mass of each gas measured at the specified conditions. (Hint: See Sample Problem 11-6.)
   a. 0.650 g occupying 1.12 L at 280 K and 1.14 atm
   b. 1.05 g occupying 2.35 L at 35°C and 0.840 atm
   c. 0.432 g occupying 7.50 x 10⁻³ mL at –23°C and 1.03 atm

21. If the density of an unknown gas is 3.20 g/L at –18°C and 2.17 atm, what is the molar mass of this gas?

22. One method of estimating the temperature of the center of the sun is based on the assumption that the center consists of gases that have an average molar mass of 2.00 g/mol. If the density of the center of the sun is 1.40 g/cm³ at a pressure of 1.30 x 10¹⁶ atm, calculate the temperature in degrees Celsius.

Gas Stoichiometry

23. Carbon monoxide reacts with oxygen to produce carbon dioxide. If 1.0 L of carbon monoxide reacts with oxygen:
   a. how many liters of oxygen are required? (Hint: See Sample Problem 11-7.)
   b. how many liters of carbon dioxide are produced?

24. Acetylene gas, C₂H₂, undergoes combustion to produce carbon dioxide and water vapor. If 7.50 L of CO₂ are produced,
   a. how many liters of C₂H₂ are required?
   b. what volume of H₂O vapor is produced?
   c. what volume of O₂ is required?

25. If liquid carbon disulfide reacts with 4.50 x 10⁻³ L of oxygen to produce the gases carbon dioxide and sulfur dioxide, what volume of each product is produced?

26. Assume that 5.60 L of H₂ at STP react with CuO according to the following equation:
   CuO(s) + H₂(g) → Cu(s) + H₂O(g)

   Make sure the equation is balanced before beginning your calculations.
   a. How many moles of H₂ react? (Hint: See Sample Problem 11-8)
   b. How many moles of Cu are produced?
   c. How many grams of Cu are produced?

27. Solid iron(III) hydroxide decomposes to produce iron(III) oxide and water vapor. If 0.75 L of water vapor is produced at STP,
   a. how many grams of iron(III) hydroxide were used?
   b. how many grams of iron(III) oxide are produced?

28. 290 L of methane, CH₄, undergoes complete combustion at 0.961 atm and 20°C, how many liters of each product are formed?

29. If air is 20.9% oxygen by volume, how many liters of air are needed for complete combustion of 25.0 L of octane vapor, C₈H₁₈?

30. A modified Haber process for making ammonia is conducted at 550°C and 2.50 x 10⁶ atm. If 10.0 kg of nitrogen (the limiting reactant) is used and the process goes to completion, what volume of ammonia is produced?

31. When liquid nitroglycerin, C₃H₅(NO₃)₃, explodes, the products are carbon dioxide, nitrogen, oxygen, and water vapor. If 5.00 x 10⁻³ g of nitroglycerin explode at STP, what is the total volume, at STP, for all gases produced?

32. The principal source of sulfur on Earth is deposits of free sulfur occurring mainly in volcanically active regions. The sulfur was initially formed by the reaction between the two volcanic vapors SO₂ and H₂S to form H₂SO₃ and S₈(s).
   What volume of each gas, at 0.961 atm and 25°C, was needed to form a sulfur deposit of 4.50 x 10⁻³ kg on the slopes of a volcano in Hawaii?

33. A 3.25 g sample of solid calcium carbide, CaC₂, reacted with water to produce acetylene gas, C₂H₂, and aqueous calcium hydroxide. If the acetylene was collected over water at 17°C and 0.974 atm, how many milliliters of acetylene were produced?

34. Balance the following chemical equation:
   Mg(s) + O₂(g) → MgO(s)

   Then, based on the quantity of reactant or product given, determine the corresponding quantities of the specified reactants or products, assuming that the system is at STP.
   a. 22.4 L O₂ → mol O₂ → _mol MgO
   b. 11.2 L O₂ → _mol O₂ → _mol MgO
   c. 1.40 L O₂ → _mol O₂ → _mol MgO

35. Assume that 8.50 L of I₂ are produced using the following reaction that takes place at STP:
   KI(aq) + Cl₂(g) → KCl(aq) + I₂(g)

   Balance the equation before beginning your calculations.
   a. How many moles of I₂ are produced?
   b. How many moles of KI were used?
   c. How many grams of KI were used?
   d. Suppose that 6.50 x 10⁻³ mL of hydrogen gas are produced through a replacement reaction involving solid iron and sulfuric acid, H₂SO₄, at STP. How many grams of iron(II) sulfate are also produced?
   e. Methanol, CH₃OH, is made by causing carbon monoxide and hydrogen gases to react at high temperature and pressure. If 4.50 x 10⁻³ mL of CO and 825 mL of H₂ are mixed,
      a. which reactant is in excess?
      b. how much of that reactant remains after the reaction?
      c. what volume of CH₄OH is produced, assuming the same pressure?

36. Suppose that 13.5 g of Al react with HCl according to the following equation, at STP:
   Al(s) + HCl(aq) → AlCl₃(aq) + H₂(g)
   Remember to balance the equation first.
   a. How many moles of Al react?
   b. How many moles of H₂ are produced?
   c. How many liters of H₂ at STP are produced?
   (Hint: See Sample Problem 11-9.)

Effusion and Diffusion

37. Quantitatively compare the rates of effusion for the following pairs of gases at the same temperature and pressure:
   a. hydrogen and nitrogen (Hint: See Sample Problem 11-10.)
   b. nitrogen and chlorine

38. What is the ratio of the average velocity of hydrogen molecules to that of neon atoms at the same temperature and pressure?

39. At a certain temperature and pressure, chlorine molecules have an average velocity of 0.0580 m/s. What is the average velocity of sulfur dioxide molecules under the same conditions?

40. A sample of helium effuses through a porous container 6.50 times faster than does unknown gas X. What is the molar mass of the unknown gas?

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43. An unknown gas effuses at 0.850 times the effusion rate of nitrogen dioxide, N₂O₂. Estimate the molar mass of the unknown gas.

44. Use the ideal gas law, PV = nRT, to derive Boyle’s law and Charles’ law.

45. A container holds 265 mL of chlorine gas, Cl₂. Assuming that the gas sample is at STP, what is its mass?

46. Suppose that 3.11 mol of carbon dioxide is at a pressure of 0.850 atm and a temperature of 39°C. What is the volume of the sample, in liters?

47. Compare the rates of diffusion of carbon monoxide, CO, and sulfur trioxide, SO₃.

48. A gas sample that has a mass of 0.993 g occupies 0.570 L. Given that the temperature is 281 K and the pressure is 1.44 atm, what is the molar mass of the gas?

49. The density of a gas is 3.07 g/L at STP. Calculate the gas’s molar mass.

50. How many moles of helium gas would it take to fill a gas balloon with a volume of 1000 cm³ when the temperature is 32°C and the atmospheric pressure is 752 mm Hg?
51. A gas sample is collected at 16°C and 0.982 atm. If the sample has a mass of 7.40 g and a volume of 3.96 L, find the volume of the gas at STP and the molar mass.

**CRITICAL THINKING**

52. Evaluating Methods  In solving a problem, what types of conditions involving temperature, pressure, volume, or number of moles would allow you to use a. the combined gas law? b. the ideal gas law?

53. Relating Ideas  Write expressions relating the rates of effusion, molar masses, and densities of two different gases, A and B.

54. Evaluating Ideas  Gay-Lussac’s law of combining volumes holds true for relative volumes at any proportionate size. Use Avogadro’s law to explain why this proportionality exists.

55. Designing Experiments  Design an experiment to prove that the proportionality described in item 54 exists.

56. Interpreting Concepts  The diagrams that follow represent equal volumes of four different gases.

![Diagrams of gases](Image)

Use the diagrams to answer the following questions:

a. Are these gases at the same temperature and pressure? How do you know? b. If the molar mass of gas B is 38 g/mol and that of gas C is 46 g/mol, which gas sample is more dense? c. To make the densities of gas samples B and C equal, which gas should expand in volume? d. If the densities of gas samples A and C are equal, what is the relationship between their molecular masses?

**TECHNOLOGY & LEARNING**

57. Graphing Calculator  Calculating Pressure Using the Ideal Gas Law

The graphing calculator can run a program that calculates the pressure in atmospheres, given the number of moles of a gas (n), volume (V), and temperature (T). Given a 0.50 mol gas sample with a volume of 19. L at 298 K, you can calculate the pressure according to the ideal gas law. Begin by using the program to carry out the calculation. Next, use it to make calculations.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data, answer these questions.

Note: Answers are written with five significant figures.

a. What is the pressure for a gas with an amount of 1.3 mol, volume of 8.0 L, and temperature of 293 K? b. What is the pressure for a gas with an amount of 2.7 mol, volume of 8.5 L, and temperature of 310 K? c. A gas with an amount of 0.75 mol and a volume of 6.0 L is measured at two different temperatures: 300 K and 275 K. At which temperature is the pressure greater?

**RESEARCH & WRITING**

58. Most elements from Groups 1, 2, and 13 will react with water, acids, or bases to produce hydrogen gas. Review the common reactions information in the Elements Handbook and answer the following:

a. Write the equation for the reaction of barium with water.

59. Group 1 metals react with oxygen to produce oxides, peroxides, or superoxides. Review the equations for these common reactions in the Elements Handbook and answer the following:

a. How do oxides, peroxides, and superoxides differ? b. What mass of product will be formed from a reaction of 5.00 L of oxygen with excess sodium? The reaction occurs at 27°C and 1 atm.

60. Most metals react with chlorine from Group 17 to produce salts. Review these common reactions in the Elements Handbook and answer the following:

a. Write the equation for the reaction of iron and chlorine gas. b. What mass of iron salt is produced if an excess of iron reacts with 450 mL of chlorine gas at 27°C and 1 atm?

**ALTERNATIVE ASSESSMENT**

61. How do scuba divers use the laws and principles that describe the behavior of gases to their advantage? What precautions do they take to prevent the bends?

62. Explain the processes involved in the liquefaction of gases. What substances that are gases under normal room conditions are typically used in the liquid form? Why?

63. Research the relationship between explosions and the establishment of Nobel Prizes. Prepare a report that describes your findings.

64. Write a summary describing how Gay-Lussac’s work on combining volumes relates to Avogadro’s study of gases. Explain how certain conclusions about gases followed logically from consideration of the work of both scientists.

65. During a typical day, record every instance in which you encounter the diffusion or effusion of gases (for example, smelling perfume).

66. Performance Qualitatively compare the molecular masses of various gases by noting how long it takes you to smell them from a fixed distance. Work only with materials that are not dangerous, such as flavor extracts, fruit peels, and onions.

67. Performance Design an experiment to gather data to verify the Ideal Gas Law. If your teacher approves of your plan, carry it out. Illustrate your data with a graph, and determine if the data are consistent with the Ideal Gas Law.