4

BIG IDEA 2: GASES

Big Idea 2

Chemical and physical properties of materials can be explained by the structure and the arrangement of atoms, ions, or molecules and the forces between them.

In this section, the laws describing gas behavior are explored. Mathematical laws relating the properties of pressure, volume, temperature, and moles of gas are considered. Densities and molar masses of ideal gases will be determined from the ideal gas law. Stoichiometric calculations will be performed for reactions involving gases. Pressures of gases in a mixture will be examined. The behavior of ideal gases will be explained by the kinetic molecular theory. Real gases will be compared to ideal gases.

Questions about gas behavior and gas laws can be either qualitative or quantitative in the multiple-choice section.

The free-response portion of the exam can include questions about an experiment involving gases such as the determination of the molar mass or molar volume of a gas. Essay questions about ideal or real gas behavior or the kinetic-molecular theory may also appear.

You should be able to

- Perform calculations with gas laws: Boyle's, Charles', Avogadro's, and ideal.
- Perform calculations with the ideal gas law to find the density or molar mass of the gas.

- Interpret or draw graphical relationships between gas variables.
- Perform stoichiometric calculations for reactions which involve gases as reactants, products, or both.
- Perform calculations with molar volume.
- Perform calculations with Dalton's law of partial pressures for a mixture of gases.
- Perform calculations for gases collected over water.
- Use kinetic-molecular theory (KMT) and knowledge of intermolecular forces to predict ideal and nonideal gas properties
- Draw representations of gas-phase materials that illustrate the intermolecular forces and the effects of changing physical conditions on the materials.
- Use data regarding real gases to identify deviations from ideal gas behavior and identify the molecular interactions leading to these deviations.

AP Tip

Mathematical equations will be provided in the free-response section of the exam, but not for the multiple-choice part. It is best to memorize the equations or know how to derive the equations from one equation.

GAS LAWS

(*Chemistry* 8th ed. pages 183–194/9th ed. pages 192–203)

The table below summarizes the gas laws that can be derived from the ideal gas law. A typical problem involves solving for the missing variable when given a set of initial and final conditions for a gas sample. **Be sure to perform all gas calculations involving temperature using the Kelvin scale, not degrees Celsius.** Temperature (K) = °C + 273.15.

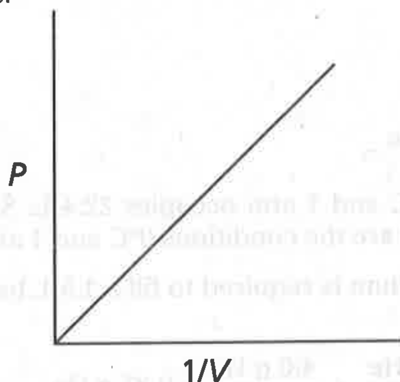
Gas Law	Equations	Definition
Boyle's law	$P = (nRT)\ 1/V$ $P_1V_1 = P_{22}$	The product of the pressure and the volume is a constant, *k*, for a trapped sample of gas at constant temperature.
Charles' law	$V = (nR/P)T$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$	The volume of a gas at constant pressure increases linearly with the temperature of the gas.

Gas Law	Equations	Definition
Avogadro's law	$V = (RT/P)n$ $\frac{V_1}{n_1} = \frac{V_2}{n_2}$	Equal volumes of gases at the same temperature and pressure contain the same number of moles of gas.
Ideal gas law	$PV = nRT$ R = universal gas constant = 0.08206 L × atm/(K × mol)	An equation of state for a gas, where the state of a gas is its condition at a given time.
Dalton's law	$P_{total} = P_1 + P_2 + P_3 + \ldots$ $P_{total} = n_{tot}RT/V$	For a mixture of gases, the sum of the pressures of the individual gases is equal to the total pressure. The total pressure depends on the total moles of gas present.

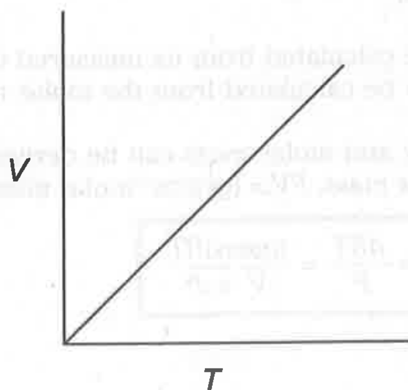
GRAPHICAL RELATIONSHIPS FOR GAS LAWS

(*Chemistry* 8th ed. pages 183–187/9th ed. pages 192–196)

BOYLE'S LAW Pressure is inversely related to volume. A graph of P vs. $1/V$ is a straight line.



CHARLES' LAW Volume is directly proportional to temperature. A graph of V vs. T is linear. Extrapolation of this graph for all gases to zero volume results in the same temperature, –273.15°C (0 K), which is absolute zero.



EXAMPLE: A sample tube containing 103.6 mL of CO gas at 20.6 torr is connected to an evacuated 1.13-L flask. What will the pressure be when the CO is allowed into the flask?

SOLUTION: $P_1V_1 = P_2V_2;\ P_2 = \frac{P_1V_1}{V_2} = \frac{(20.6\text{ torr})(0.1036\text{ L})}{1.47\text{ L}} = 1.45\text{ torr}$

Note: Units of volume must cancel out. Remember to convert to the same units for all volume measurements.

EXAMPLE: A quantity of gas at 27.0°C is heated in a closed vessel until the pressure is doubled. To what temperature is the gas heated?

SOLUTION: Pressure is directly proportional to temperature in Kelvin units, so before you do any computations, convert the temperature to Kelvin. If the pressure is doubled, the temperature will also double. 27.0°C + 273.15 = 300.2 K. The answer is 600.4 K (or 327.2°C).

GAS STOICHIOMETRY

(*Chemistry* 8th ed. pages 194–199/9th ed. pages 203–208)

If the pressure, volume, and temperature of ideal gases are known, stoichiometric calculations can be performed. This is discussed in more detail in Big Idea 3.

MOLAR VOLUME

(*Chemistry* 8th ed. pages 194–195/9th ed. pages 203–204)

One mole of an ideal gas at 0°C and 1 atm occupies 22.4 L. Standard temperature and pressure (STP) are the conditions 0°C and 1 atm.

EXAMPLE: What mass of helium is required to fill a 1.5-L balloon at STP?

SOLUTION: $1.5\text{ L} \times \frac{1\text{ mol He}}{22.4\text{ L}} \times \frac{4.0\text{ g He}}{1\text{ mol}} = 0.27\text{ g He}$

MOLAR MASS OF A GAS

(*Chemistry* 8th ed. pages 198–199/9th ed. pages 207–208)

The molar mass of a gas can be calculated from its measured density, *d*. The density of a gas can also be calculated from the molar mass of the gas.

The equation relating density and molar mass can be derived from $PV = nRT$. Since n = grams/molar mass, PV = (grams /molar mass) RT.

$$\text{Molar Mass} = \frac{dRT}{P} = \frac{(\text{mass})RT}{V \times P}$$

EXAMPLE: A sample of gas weighing 0.800 g occupies a 256-mL flask at 100°C and 750.0 torr. Determine the molar mass of the gas.

SOLUTION: First, convert pressure from torr to atm.

$$750.0 \text{ torr} \times \frac{1.00 \text{ atm}}{760.0 \text{ torr}} = 0.986 \text{ atm}$$

$$\text{Molar mass} = \frac{0.800 \text{ g} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 373 \text{ K}}{0.256 \text{ L} \times 0.986 \text{ atm}} = 97.0 \text{ g/mol}$$

DALTON'S LAW

(*Chemistry* 8th ed. pages 199–205/9th ed. pages 208–214)

For a mixture of gases in a container, the total pressure is the sum of the pressures that each gas would exert if it were alone.

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

The partial pressure of a gas, P_{gas}, is the pressure that a particular gas would exert if it were alone in the container. The subscripts refer to the individual gases (gas_1, gas_2, and so on). The partial pressure of each gas can be calculated from the ideal gas law:

$$P_1 = n_1RT/V$$

The total pressure of the mixture, P, can be represented as

$$P_{total} = n_{tot}RT/V$$

The mole fraction is the ratio of the number of moles of a given component in a mixture to the total number of moles in the mixture.

$$X_1 = n_1/n_{tot}$$

EXAMPLE: A mixture of 1.00 g of H_2 and 1.00 g of He is placed in a 1.00-L container at 27°C. Calculate the mole fraction and partial pressure of each gas. Calculate the total pressure in the container.

SOLUTION: $n_{H_2} = 1.00 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.02 \text{ g } H_2} = 0.496 \text{ mol } H_2$

$$n_{He} = 1.00 \text{ g He} \times \frac{1 \text{ mol He}}{4.00 \text{ g}} = 0.250 \text{ mol He}$$

$$X_{H2} = \frac{0.496 \text{ mol}}{0.496 \text{ mol} + 0.250 \text{ mol}} = 0.665; \quad X_{He} = \frac{0.250 \text{ mol}}{0.496 \text{ mol} + 0.250 \text{ mol}} = 0.335$$

$$P_{H_2} = \frac{n_{H_2}RT}{V} = \frac{0.496 \text{ mol} \times \dfrac{0.08206 \text{ L} \times \text{atm}}{\text{mol} \times \text{K}} \times 300 \text{ K}}{1.00 \text{ l}} = 12.2 \text{ atm}$$

$$P_{He} = \frac{n_{He}RT}{V} = \frac{0.250 \text{ mol} \times \dfrac{0.08206 \text{ L} \times \text{atm}}{\text{mol} \times \text{K}} \times 300 \text{ K}}{1.0 \text{ L}} = 6.15 \text{ atm}$$

$$P_{total} = P_{He} + P_{H_2} = 12.2 \text{ atm} + 6.15 \text{ atm} = 18.4 \text{ atm}$$

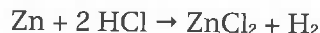
GAS COLLECTION OVER WATER

(*Chemistry* 8th ed. pages 202–205/9th ed. pages 211–214)

A gas can be collected by displacement of water. A mixture of gases results due to a mixture of water vapor and the gas being collected. Refer to Figure 5.13 on page 204 of the 8th edition and page 213 of the 9th edition of *Chemistry*.

EXAMPLE: A sample weighing 0.986 g contains zinc and some impurities. Excess hydrochloric acid is added and reacts with the zinc but not the impurities. Determine the percentage of zinc in the sample if 240.0 mL of hydrogen gas is collected over water at 30.0°C and 1.032 atm. The vapor pressure of water at this temperature is 0.042 atm.

SOLUTION: Beginning with the balanced equation:



$Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

One mole of zinc will produce one mole of hydrogen gas.

Using Dalton's law and the vapor pressure of water at the specified temperature:

$P_{total} = P_{atm} = P_{H_2} + P_{H_2O}$; $P_{H_2} = P_{atm} - P_{H_2O}$

$P_{H_2} = 1.032 \text{ atm} - 0.042 \text{ atm} = 0.990 \text{ atm}$

$$n_{H_2} = \frac{P_{H_2} \times V}{RT} = \frac{0.990 \text{ atm} \times 0.240 \text{ L}}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 303.2 \text{ K}} = 0.00955 \text{ mol } H_2$$

$$0.00955 \text{ mol } H_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } H_2} \times \frac{65.4 \text{ g Zn}}{1 \text{ mol Zn}} = 0.625 \text{ g Zn}$$

$$\%Zn = \frac{\text{g Zn} \times 100\%}{\text{g sample}} = \frac{0.625 \text{ g}}{0.986 \text{ g}} \times 100\% = 63.3\% Zn$$

Note that the vapor pressure of water at various temperatures can be found in tables of known physical constants. When you solve this kind of problem, the vapor pressure of water will be provided.

KINETIC-MOLECULAR THEORY

(*Chemistry* 8th ed. pages 205–213/9th ed. pages 214–223)

The kinetic-molecular theory is a model that attempts to explain the properties of an ideal gas. The postulates of the kinetic-molecular theory are

1. The volume of the individual particles of a gas can be assumed to be negligible.
2. The gas particles are in constant motion. The pressure exerted by the gas is due to the collisions of the gases with the walls of the container.
3. The gases are not attracted to one another.
4. The average kinetic energy of a gas is directly proportional to the Kelvin temperature.

$$(KE)_{average} = 3/2\ RT = ½\ mv^2_{avg}$$

Note that the value of R here is 8.31 J mol^{-1} K^{-1}. The other values of R do not have the correct units for energy calculations.

EXAMPLE: Three identical flasks are filled with three different gases.

Flask A: CO at 760 torr and 0°C

Flask B: N_2 at 250 torr and 0°C

Flask C: H_2 at 100 torr and 0°C

In which flask will the molecules have the greatest average kinetic energy? In which flask will the molecules have the greatest average velocity?

SOLUTION: All molecules will have the same average kinetic energy since they are all at the same temperature. Flask C will have the greatest average velocity since hydrogen has the lowest molar mass. At constant T, the lightest molecules are fastest, on average.

REAL GASES

(*Chemistry* 8th ed. pages 214–217/9th ed. pages 224–227)

The ideal gas model fails at high pressure and low temperature. Real gas molecules take up space and experience attractive forces between molecules. At high pressure there is less empty space between molecules, and the volume of molecules becomes more significant. An ideal gas could be compressed to zero volume, but for a real gas, as the pressure doubles, the volume of empty space cannot continue to be halved. As the temperature decreases, the molecules have less kinetic energy to overcome the attractive forces between gas molecules; these attractive forces may cause the gas to condense.

The van der Waals equation modifies the assumptions of the kinetic-molecular theory to fit the behavior of real gases. The van der Waals constant a adjusts for the presence of intermolecular attractions; the van der Waals constant b adjusts for the size of the atom or molecule.

$$[P_{obs} + a(n/V)^2] \times (V - nb) = nRT$$

EXAMPLE: Which of the following gases would you expect to have the largest value of the van der Waals constant, b: H_2, N_2, CH_4, C_2H_6, or C_3H_8?

SOLUTION: Since the van der Waals constant, b, is a measure of the size of the molecule. C_3H_8 has the largest molar volume and should have the largest value of b.

EXAMPLE: Which of the following gases would you expect to have the largest value of the van der Waals constant, a: H_2, CO_2, CH_4, or N_2?

SOLUTION: CO_2 has the largest value for a, which measures intermolecular attractions. All the molecules are nonpolar so the only force present is an induced dipole or London force which increases as the number of electrons and protons in the molecule increases.

MULTIPLE-CHOICE QUESTIONS

No calculators are to be used in this section.

1. Under what conditions does a gas behave more like a real gas than an ideal gas?
   (A) high temperature and low pressure
   (B) high temperature and high pressure
   (C) low temperature and low pressure
   (D) low temperature and high pressure

2. What is the volume of 3.00 mol of gas at STP?
   (A) 22.4 L
   (B) 3 × 22.4 L
   (C) 3 × 22.4 L × 273/760
   (D) It cannot be determined without knowing which gas is involved.

3. An ideal gas of volume 189. mL is collected over water at 30°C and 777 torr. The vapor pressure of water is 32 torr at 30°C. What pressure is exerted by the dry gas under these conditions?
   (A) 320 torr
   (B) 745 torr
   (C) 777 torr
   (D) 32/777 torr

4. A 14.0-L cylinder contains 5.60 g N_2, 40.0 g Ar, and 6.40 g O_2. What is the total pressure in atm at 27°C? (R = the ideal gas constant.)
   (A) 26 R
   (B) 30 R
   (C) 60 R
   (D) 120 R

5. In a closed rigid system, 7.0 mol CO_2, 7.0 mol Ar, 7.0 mol N_2, and 4.0 mol Ne are trapped, with a total pressure of 10.0 atm. What is the partial pressure exerted by the neon gas?
(A) 1.6 atm
(B) 4.0 atm
(C) 10.0 atm
(D) 21.0 atm

6. Consider the reaction:

$C_2H_6(g) + 7/2\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

If 6.00 g ethane, $C_2H_6(g)$ burn, what volume of $CO_2(g)$ will be formed at STP?
(A) 0.200 L
(B) 2.20 L
(C) 9.00 L
(D) 22.4 L

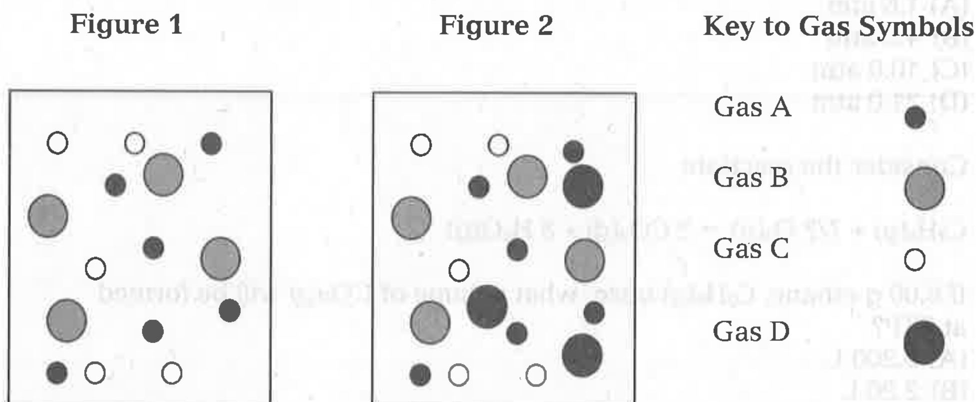
7. Cl_2 and F_2 combine to form a gaseous product; one volume of Cl_2 reacts with three volumes of F_2 yielding two volumes of product. Assuming constant conditions of temperature and pressure, what is the formula of the product?
(A) Cl_2F_2
(B) ClF_2
(C) Cl_2F
(D) ClF_3

8. Decreasing the temperature of an ideal gas from 80°C to 40°C causes the average kinetic energy to
(A) decrease by a factor of two
(B) increase by a factor of two
(C) increase by a factor of four
(D) decrease by less than a factor of two

9. The average speed of the molecules of a gas is proportional to the
(A) reciprocal of absolute temperature, $(1/T)$
(B) absolute temperature
(C) square root of the absolute temperature
(D) square of the absolute temperature

10. A 5.00-L vessel contains 2.00 moles of helium and 3.00 moles of hydrogen at a pressure of 10.0 atm. Maintaining a constant temperature, an additional 3.00 moles of hydrogen are added. What is the partial pressure of hydrogen gas in the vessel at the end? (Assume that the gases behave ideally.)
(A) 6.00 atm
(B) 10.0 atm
(C) 12.0 atm
(D) 20.0 atm

Questions 11–13 refer to Figure 1 and Figure 2 below. On the right is a key to identify the gases in the problems.



11. In Figure 1 there are three gases present, Gas A, B, and C. The number of spheres represents the number of moles of each gas present. If the total pressure within the vessel is 5.00 atm, what is the partial pressure of Gas B?
(A) 1.00 atm
(B) 1.33 atm
(C) 1.50 atm
(D) 1.67 atm

12. Suppose that another gas D is added to the vessel in Figure 1, keeping the temperature and the volume constant. This is represented by Figure 2 above. What happens to the partial pressure of Gas B?
(A) increases
(B) decreases
(C) stays the same
(D) cannot be predicted

13. What happens to the total pressure for the system described in Question 12? (In other words, compare the total pressure in Figure 1 versus Figure 2.)
(A) increases
(B) decreases
(C) remains the same
(D) cannot be predicted

14. Increasing the pressure on a gas in a rigid container at constant temperature will
(A) increase the average kinetic energy of the gas molecules
(B) increase the influence of intermolecular attractions, because the molecules will be closer together
(C) increase the influence of intermolecular attractions, because the molecules will be moving more slowly
(D) decrease the average kinetic energy of the gas molecules

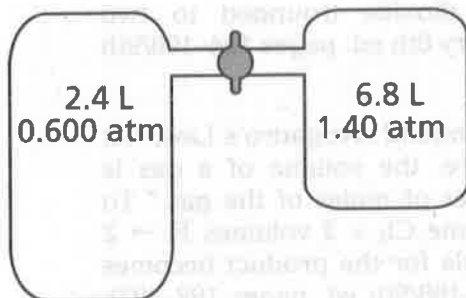
15. Consider two samples of a gas in identical rigid containers, Sample A at 273 K and Sample B at 300 K. Which of the following statements is true?
(A) The average kinetic energy of Sample A is less than that of Sample B.
(B) The samples have the same average velocity.
(C) The samples must contain the same number of moles of gas.
(D) The pressure of Sample A is more than the pressure of Sample B.

FREE-RESPONSE QUESTIONS

Calculators may be used for this section.

1. (a) Compare the temperature of freshly made coffee made at lower altitudes to coffee made at higher altitudes. Cite the relevant intermolecular forces in your explanation.
   (b) Under what conditions are gases most "ideal"? Explain why, in terms of KMT and intermolecular forces of attraction.
   (c) One mole of water and one mole of propane are placed in separate, closed, 1-L containers and heated to 110°C. The pressure in the water vapor container is less than the pressure in the propane container. Explain this observation.
   (d) How does gaseous pressure relate to changes in volume? Explain.

2. Assume that two cylinders at 27°C are connected by a closed stopcock (valve) system. The right-hand cylinder contains 2.40 L of hydrogen at 0.600 atm; the left cylinder is larger and contains 6.80 L of helium at 1.40 atm.



   (a) How many moles of each gas are present?
   (b) What is the total pressure when the valve is open?
   (c) Determine the partial pressure of these two gases at 27.0°C when the stopcock is opened.

Answers

MULTIPLE-CHOICE QUESTIONS

1. D At low temperature and high pressure, the molecules are closer together and therefore the forces between molecules become more important as they are stronger. The actual (finite) volume of individual molecules also becomes more important

as more of the total space is actually occupied by finite molecular volume (*Chemistry* 8th ed. pages 214–217/9th ed. pages 224–227). LO 2.4

2. **B** The molar volume of all gases at STP is 22.4 L, so three moles would occupy 22.4 L/mol × 3 moles (*Chemistry* 8th ed. pages 214–217/9th ed. pages 224–227). LO 2.6

3. **B** The total pressure = 777 torr. Of this, 32 torr is due to the water vapor, hence 777 – 32 = 745 torr of pressure are allocated to the dry gas (*Chemistry* 8th ed. pages 199–205, 471–478/9th ed. pages 208–214, 483–490). LO 2.6

4. **B** Use $P_{Total} = N_{Total}RT/V$ to determine the pressure. From the mass of each of the gases you can find 0.200 mol of N_2, 1.00 mol of Ar, and 0.200 mol of O_2 to give a total number of moles (n) of 1.4 mol. Therefore $P_{Total} = R \times 1.40 \text{ mol} \times 300 \text{ K} / 14.0 \text{ L} = 30 R$. Note: You do not need a calculator to divide 1.4 by 14 and then multiply by 300! (*Chemistry* 8th ed. pages 188–194/9th ed. pages 197–203). LO 2.6

5. **A** The total number of moles is 25.0 (7.0 + 7.0 + 7.0 + 4.0 = 25.0), hence the Ne is 4.0/25.0 of the total amount of gas and exerts 4.0/25.0 of the total pressure, or 4.0/25.0 × 10.0 atm = 1.6 atm (*Chemistry* 8th ed. pages 199–205/9th ed. pages 208–214). LO 2.6

6. **C** 6.0 g of ethane / 30. g/mol yields 0.20 mol of ethane, which forms twice that number of moles of carbon dioxide. Since one mol of gas occupies 22.4 L at STP, 22.4 L/mol × 0.20 mol × 2 CO_2/C_2H_6 = 8.96 L of carbon dioxide (rounded to two significant figures = 9.0 L) (*Chemistry* 8th ed. pages 194–199/9th ed. pages 203–208). LO 2.6

7. **D** This problem assumes you understand Avogadro's Law, "At constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of the gas." To apply that to this problem, 1 volume Cl_2 + 3 volumes $F_2 \rightarrow$ 2 volumes Cl_xF_{3x}. The simplest formula for the product becomes ClF_3 (*Chemistry* 8th ed. pages 183–188/9th ed. pages 192–197). LO 2.6

8. **D** Be careful here to note the difference between the kinds of temperature scales and what they mean. Even though the Celsius temperature is half as much, the average kinetic energy is proportional to the Kelvin temperature. In this case that ratio is only 353/313 = 1.13 so the average kinetic energy decreases only by a factor of 1.13 (*Chemistry* 8th ed. pages 203–212/9th ed. pages 212–222). LO 2.6

9. **C** In this case the question is about the average speed (not the energy) of the molecules. Review root-mean-square velocity (*Chemistry* 8th ed. pages 211–212/9th ed. pages 221–222). LO 2.6

10. D The initial partial pressure of the hydrogen = 10 atm × (3.0/5.0) = 6.0 atm. The amount of hydrogen present is doubled from 3.0 moles to 6.0 moles, therefore the partial pressure will increase proportionally to 12.0 atm. According to KMT, the volume of the particles is negligible and the particles do not interact with each other, so it does not matter what other gases are added or what other gases are present (*Chemistry* 8th ed. pages 199–203/9th ed. pages 208–212). LO 2.4

11. B The partial pressure = mole fraction × total pressure. In this case, partial pressure of Gas B = (4/15)(5.00 atm) = 1.33 atm (*Chemistry* 8th ed. pages 199–203/9th ed. pages 208–212). LO 2.4

12. C Since the amount of Gas B remains the same, there would be no change in the partial pressure of this gas (*Chemistry* 8th ed. pages 199–203/9th ed. pages 208–212). LO 2.4

13. A Since the total number of moles of gas in the vessel has now increased, the total pressure will also increase (*Chemistry* 8th ed. pages 199–203/9th ed. pages 208–212). LO 2.4

14. B Intermolecular forces are more significant when molecules are closer together (*Chemistry* 8th ed. pages 214–216/9th ed. pages 224–226). LO 2.4

15. A According to KMT, temperature is a measure of the kinetic energy of the gas, which in turn is related to the average velocity. A higher temperature means higher average velocity (*Chemistry* 8th ed. page 210/9th ed. page 220). LO 2.4

FREE-RESPONSE QUESTIONS

1. (a) At low altitudes, the amount of air above the surface of the earth, and therefore the total atmospheric pressure, would be greater. Water boils at a higher temperature under such conditions since a higher vapor pressure is required for the liquid to become a gas. The coffee made at higher altitude will not be as hot (*Chemistry* 8th ed. pages 481–482/9th ed. pages 493–494). LO 2.16

   (b) In this question, be careful not to confuse "conditions" with "physical properties" or "characteristics of the gaseous molecules." The conditions under which real gases are most ideal are those of low pressure and high temperature. "Low pressure" implies that the gas molecules are far apart from each other, so the forces of attraction between the molecules are weak. The finite volume of individual molecules is also, therefore, a small part of the total volume occupied by the gas, as required by KMT. Finally, at higher temperatures the molecules are moving so rapidly that the effect of intermolecular attractions is negligible (*Chemistry* 8th ed. pages 188–194, 214–217/9th ed. pages 197–203, 224–227). LO 2.4

(c) Water molecules are polar and attracted to each other by hydrogen bonding. At temperatures just above water's boiling point, these attractions will be more important. Because of this attraction, the molecules will hit the container walls less frequently and the pressure will be less than predicted by the ideal gas law. Nonpolar propane molecules only experience London dispersion forces of attraction which are weak and easily overcome (*Chemistry* 8th ed. pages 214–217/9th ed. pages 224–227). LO 2.4

(d) Gaseous pressure is inversely proportional to the volume of the container. If only the volume of the container is less, for example, the molecules have less room to move around before striking the sides and therefore strike the sides of the container more often. This assumes that the temperature is held constant (*Chemistry* 8th ed. pages 183–186/9th ed. pages 192–195). LO 2.4

Note: What assumption is made in your answer to D?

2. Solving this problem has three steps. (1) To determine the pressure exerted by each gas, first calculate the number of moles of each ($n = PV / RT$), (2) and then use the total number of moles of gas and $p = nRT/V$ to calculate the total pressure when the total volume is 9.20 L. (3) Finally, the partial pressure of each gas is the total pressure × the mol fraction of that gas (*Chemistry* 8th ed. pages 188–194 and 199–200/9th ed. pages 197–203 and 208–209). LO 2.6

For hydrogen: (0.600 atm × 2.40 L) / (R × 300. K) = 0.0585 mol of H_2

For helium: (1.40 atm × 6.80 L) / (R × 300. K) = 0.397 mol of He

Total number of moles = 0.0585 + 0.397 mol = 0.456 mol

Total pressure is then (0.456 mol × R × 300. K) / (9.20 L) = 1.22 atm

Use the mol fractions: (0.0585/0.456) × 1.22 atm = 0.157 atm H_2

(0.397/0.456) × 1.22 atm = 1.06 atm He