6

Gases

HOW OFTEN DO GASES APPEAR ON THE EXAM?

In the multiple-choice section, this topic appears in about 5 out of 75 questions. In the free-response section, this topic appears almost every year.

STANDARD TEMPERATURE AND PRESSURE (STP)

You should be familiar with standard temperature and pressure (STP), which comes up fairly often in problems involving gases.

At STP:	Pressure = 1 atmosphere = 760 millimeters of mercury (mmHg)
	Temperature = 0°C = 273 K
At STP, 1 mole of gas occupies 22.4 liters	

KINETIC MOLECULAR THEORY

Most of the gas problems you will see on the test will assume that gases behave in what is called an ideal manner. For ideal gases, the following assumptions can be made:

- The kinetic energy of an ideal gas is directly proportional to its absolute temperature: The greater the temperature, the greater the average kinetic energy of the gas molecules.

The Total Kinetic Energy of a Gas Sample
$$KE = \frac{3}{2}nRT$$ R = the gas constant; 8.31 joules/mol-K T = absolute temperature (K) n = number of moles (mol)

The Average Kinetic Energy of a Single Gas Molecule
$$KE = \frac{1}{2}mv^2$$ m = mass of the molecule (kg) v = speed of the molecule (meters/sec) KE is measured in joules

- If several different gases are present in a sample at a given temperature, all the gases will have the same average kinetic energy. That is, the average kinetic energy of a gas depends only on the absolute temperature, not on the identity of the gas.
- The volume of an ideal gas particle is insignificant when compared with the volume in which the gas is contained.
- There are no forces of attraction between the gas molecules in an ideal gas.
- Gas molecules are in constant motion, colliding with one another and with the walls of their container.

THE IDEAL GAS EQUATION

You can use the ideal gas equation to calculate any of the four variables relating to the gas, provided that you already know the other three.

The Ideal Gas Equation
$$PV = nRT$$ P = the pressure of the gas (atm) V = the volume of the gas (L) n = the number of moles of gas T = the absolute temperature of the gas (K) R = the gas constant, 0.0821 L-atm/mol-K

You can also manipulate the ideal gas equation to figure out how changes in each of its variables affect the other variables.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P = the pressure of the gas (atm)
V = the volume of the gas (L)
T = the absolute temperature of the gas (K)

You should be comfortable with the following simple relationships:

- If the volume is constant: As pressure increases, temperature increases; as temperature increases, pressure increases.
- If the temperature is constant: As pressure increases, volume decreases; as volume increases, pressure decreases. That's Boyle's law.
- If the pressure is constant: As temperature increases, volume increases; as volume increases, temperature increases. That's Charles's law.

DALTON'S LAW

Dalton's law states that the total pressure of a mixture of gases is just the sum of all the partial pressures of the individual gases in the mixture.

Dalton's Law
$P_{total} = P_a + P_b + P_c + ...$

You should also note that the partial pressure of a gas is directly proportional to the number of moles of that gas present in the mixture. So if 25 percent of the gas in a mixture is helium, then the partial pressure due to helium will be 25 percent of the total pressure.

Partial Pressure
$P_a = (P_{total})(X_a)$
$X_a = \dfrac{\text{moles of gas A}}{\text{total moles of gas}}$

GRAHAM'S LAW

Part of the first assumption of kinetic molecular theory was that all gases at the same temperature have the same average kinetic energy. Knowing this, we can find the average speed of a gas molecule at a given temperature.

$$u_{rms} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{M}}$$

u_{rms} = average speed of a gas molecule (meters/sec)
T = absolute temperature (K)
m = mass of the gas molecule (kg)
M = molecular weight of the gas (kg/mol)
k = Boltzmann's constant, 1.38×10^{-23} joule/K
R = the gas constant, 8.31 joules/mol-K

By the way, you may have noticed that Boltzmann's constant, k, and the gas constant, R, differ only by a factor of Avogadro's number, N_A, the number of molecules in a mole. That is, $R = kN_A$.

Knowing that the average kinetic energy of a gas molecule is dependent only on the temperature, we can compare the average speeds (and the rates of effusion) of two different gases in a sample. The equation used to do this is called Graham's law.

Graham's Law

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

r = rate of effusion of a gas or average speed of the molecules of a gas
M = molecular weight

You should note that Graham's law states that at a given temperature, lighter molecules move faster than heavier molecules.

VAN DER WAALS EQUATION

At low temperature and/or high pressure, gases behave in a less-than-ideal manner. That's because the assumptions made in kinetic molecular theory become invalid under conditions where gas molecules are packed too tightly together.

Two things happen when gas molecules are packed too tightly.

- *The volume of the gas molecules becomes significant.*
  The ideal gas equation does not take the volume of gas molecules into account, so the actual volume of a gas under nonideal conditions will be larger than the volume predicted by the ideal gas equation.

- *Gas molecules attract one another and stick together.*
  The ideal gas equation assumes that gas molecules never stick together. When a gas is packed tightly together, van der Waals forces (dipole–dipole attractions and London dispersion forces) become significant, causing some gas molecules to stick together. When gas molecules stick together, there are fewer particles bouncing around and creating pressure, so the real pressure in a nonideal situation will be smaller than the pressure predicted by the ideal gas equation.

The van der Waals equation adjusts the ideal gas equation to take nonideal conditions into account.

Van der Waals Equation

$$(P + \frac{n^2a}{V^2})(V - nb) = nRT$$

P = the pressure of the gas (atm)
V = the volume of the gas (L)
n = the number of moles of gas (mol)
T = the absolute temperature of the gas (K)
R = the gas constant, 0.0821 L-atm/mol-K
a = a constant, different for each gas, that takes into account the attractive forces between molecules
b = a constant, different for each gas, that takes into account the volume of each molecule

DENSITY

You may be asked about the density of a gas. The density of a gas is measured in the same way as the density of a liquid or solid: in mass per unit volume.

Density of a Gas

$$D = \frac{m}{V}$$

D = density
m = mass of gas, usually in grams
V = volume occupied by a gas, usually in liters

CHAPTER 6 QUESTIONS

MULTIPLE-CHOICE QUESTIONS

<u>Questions 1–5</u>

(A) H_2
(B) He
(C) O_2
(D) N_2
(E) CO_2

1. This is the most plentiful gas in the earth's atmosphere.

2. A 1 mole sample of this gas occupying 1 liter will have the greatest density.

3. At a given temperature, this gas will have the greatest rate of effusion.

4. The molecules of this nonpolar gas contain polar bonds.

5. The molecules of this gas contain triple bonds.

6. The temperature of a sample of an ideal gas confined in a 2.0 L container was raised from 27°C to 77°C. If the initial pressure of the gas was 1,200 mmHg, what was the final pressure of the gas?

(A) 300 mmHg
(B) 600 mmHg
(C) 1,400 mmHg
(D) 2,400 mmHg
(E) 3,600 mmHg

7. A sealed container containing 8.0 grams of oxygen gas and 7.0 of nitrogen gas is kept at a constant temperature and pressure. Which of the following is true?

(A) The volume occupied by oxygen is greater than the volume occupied by nitrogen.
(B) The volume occupied by oxygen is equal to the volume occupied by nitrogen.
(C) The volume occupied by nitrogen is greater than the volume occupied by oxygen.
(D) The density of nitrogen is greater than the density of oxygen.
(E) The average molecular speeds of the two gases are the same.

8. A gas sample contains 0.1 mole of oxygen and 0.4 moles of nitrogen. If the sample is at standard temperature and pressure, what is the partial pressure due to nitrogen?

(A) 0.1 atm
(B) 0.2 atm
(C) 0.5 atm
(D) 0.8 atm
(E) 1.0 atm

9. A mixture of gases contains 1.5 moles of oxygen, 3.0 moles of nitrogen, and 0.5 moles of water vapor. If the total pressure is 700 mmHg, what is the partial pressure of the nitrogen gas?

(A) 70 mmHg
(B) 210 mmHg
(C) 280 mmHg
(D) 350 mmHg
(E) 420 mmHg

10. A mixture of helium and neon gases has a total pressure of 1.2 atm. If the mixture contains twice as many moles of helium as neon, what is the partial pressure due to neon?

(A) 0.2 atm
(B) 0.3 atm
(C) 0.4 atm
(D) 0.8 atm
(E) 0.9 atm

11. Nitrogen gas was collected over water at 25°C. If the vapor pressure of water at 25°C is 23 mmHg, and the total pressure in the container is measured at 781 mmHg, what is the partial pressure of the nitrogen gas?

(A) 23 mmHg
(B) 46 mmHg
(C) 551 mmHg
(D) 735 mmHg
(E) 758 mmHg

12. When 4.0 moles of oxygen are confined in a 24-liter vessel at 176°C, the pressure is 6.0 atm. If the oxygen is allowed to expand isothermally until it occupies 36 liters, what will be the new pressure?

(A) 2 atm
(B) 3 atm
(C) 4 atm
(D) 8 atm
(E) 9 atm

13. A gas sample is confined in a 5-liter container. Which of the following will occur if the temperature of the container is increased?

I. The kinetic energy of the gas will increase.
II. The pressure of the gas will increase.
III. The density of the gas will increase.

(A) I only
(B) II only
(C) I and II only
(D) I and III only
(E) I, II, and III

14. A 22.0 gram sample of an unknown gas occupies 11.2 liters at standard temperature and pressure. Which of the following could be the identity of the gas?

(A) CO_2
(B) SO_3
(C) O_2
(D) N_2
(E) He

15. A gaseous mixture at a constant temperature contains O_2, CO_2, and He. Which of the following lists the three gases in order of increasing average molecular speeds?

(A) O_2, CO_2, He
(B) O_2, He, CO_2
(C) He, CO_2, O_2
(D) He, O_2, CO_2
(E) CO_2, O_2, He

16. Which of the following conditions would be most likely to cause the ideal gas laws to fail?

I. High pressure
II. High temperature
III. Large volume

(A) I only
(B) II only
(C) I and II only
(D) I and III only
(E) II and III only

17. Which of the following expressions is equal to the density of helium gas at standard temperature and pressure?

(A) $\frac{1}{22.4}$ g/L
(B) $\frac{2}{22.4}$ g/L
(C) $\frac{1}{4}$ g/L
(D) $\frac{4}{22.4}$ g/L
(E) $\frac{4}{4}$ g/L

18. An ideal gas is contained in a 5.0 liter chamber at a temperature of 37°C. If the gas exerts a pressure of 2.0 atm on the walls of the chamber, which of the following expressions is equal to the number of moles of the gas? The gas constant, R, is 0.08 (L-atm)/(mol-K).

(A) $\frac{(2.0)(5.0)}{(0.08)(37)}$ moles
(B) $\frac{(2.0)(0.08)}{(5.0)(37)}$ moles
(C) $\frac{(2.0)(0.08)}{(5.0)(310)}$ moles
(D) $\frac{(2.0)(310)}{(0.08)(5.0)}$ moles
(E) $\frac{(2.0)(5.0)}{(0.08)(310)}$ moles

19. Which of the following gases would be expected to have a rate of effusion that is one-third as large as that of H_2?

(A) O_2
(B) N_2
(C) He
(D) H_2O
(E) CO_2

20. In an experiment $H_2(g)$ and $O_2(g)$ were completely reacted, above the boiling point of water, according to the following equation in a sealed container of constant volume and temperature:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

If the initial pressure in the container before the reaction is denoted as P_i, which of the following expressions gives the final pressure, assuming ideal gas behavior?

(A) P_i
(B) $2\,P_i$
(C) $(3/2)P_i$
(D) $(2/3)P_i$
(E) $(1/2)P_i$

21. Nitrogen gas was collected over water at a temperature of 40°C, and the pressure of the sample was measured at 796 mmHg. If the vapor pressure of water at 40°C is 55 mmHg, what is the partial pressure of the nitrogen gas?

(A) 55 mmHg
(B) 741 mmHg
(C) 756 mmHg
(D) 796 mmHg
(E) 851 mmHg

22. A balloon occupies a volume of 1.0 liter when it contains 0.16 grams of helium at 37°C and 1 atm pressure. If helium is added to the balloon until it contains 0.80 grams while pressure and temperature are kept constant, what will be the new volume of the balloon?

(A) 0.50 liters
(B) 1.0 liters
(C) 2.0 liters
(D) 4.0 liters
(E) 5.0 liters

23. An ideal gas fills a balloon at a temperature of 27°C and 1 atm pressure. By what factor will the volume of the balloon change if the gas in the balloon is heated to 127°C at constant pressure?

(A) $\frac{27}{127}$

(B) $\frac{3}{4}$

(C) $\frac{4}{3}$

(D) $\frac{2}{1}$

(E) $\frac{127}{27}$

24. A gas sample with a mass of 10 grams occupies 6.0 liters and exerts a pressure of 2.0 atm at a temperature of 26°C. Which of the following expressions is equal to the molecular mass of the gas? The gas constant, R, is 0.08 (L-atm)/(mol-K).

(A) $\frac{(10)(0.08)(299)}{(2.0)(6.0)}$ g/mol

(B) $\frac{(299)(0.08)}{(10)(2.0)(6.0)}$ g/mol

(C) $\frac{(2.0)(6.0)(299)}{(10)(0.08)}$ g/mol

(D) $\frac{(10)(2.0)(6.0)}{(299)(0.08)}$ g/mol

(E) $\frac{(2.0)(6.0)}{(10)(299)(0.08)}$ g/mol

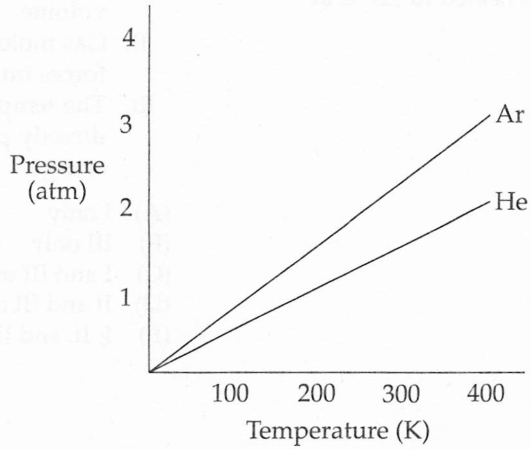
25. Which of the following assumptions is (are) valid based on kinetic molecular theory?

I. Gas molecules have negligible volume.
II. Gas molecules exert no attractive forces on one another.
III. The temperature of a gas is directly proportional to its kinetic energy.

(A) I only
(B) III only
(C) I and III only
(D) II and III only
(E) I, II, and III

PROBLEMS

1.



The graph above shows the changes in pressure with changing temperature of gas samples of helium and argon confined in a closed 2-liter vessel.

(a) What is the total pressure of the two gases in the container at a temperature of 200 K?

(b) How many moles of helium are contained in the vessel?

(c) How many molecules of helium are contained in the vessel?

(d) What is the ratio of the average speeds of the helium atoms to the average speeds of the argon atoms?

(e) If the volume of the container were reduced to 1 liter at a constant temperature of 300 K, what would be the new pressure of the helium gas?

2.

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

The reaction above took place, and 1.45 liters of oxygen gas were collected over water at a temperature of 29°C and a pressure of 755 millimeters of mercury. The vapor pressure of water at 29°C is 30.0 millimeters of mercury.

(a) What is the partial pressure of the oxygen gas collected?

(b) How many moles of oxygen gas were collected?

(c) What would be the dry volume of the oxygen gas at a pressure of 760 millimeters of mercury and a temperature of 273 K?

(d) What was the mass of the $KClO_3$ consumed in the reaction?

ESSAYS

3. Equal molar quantities of two gases, O_2 and H_2O, are confined in a closed vessel at constant temperature.

   (a) Which gas, if any, has the greater partial pressure?

   (b) Which gas, if any, has the greater density?

   (c) Which gas, if any, has the greater concentration?

   (d) Which gas, if any, has the greater average kinetic energy?

   (e) Which gas, if any, will show the greater deviation from ideal behavior?

   (f) Which gas, if any, has the greater average molecular speed?

4. SF_6, H_2O, and CO_2 are all greenhouse gases because they absorb infrared radiation in the atmosphere. As such they are of interest to many researchers.

   (a) Of the three gases, which deviates the most from ideal gas behavior?

   (b) If all three ideal gases are held at constant temperature, which gas would have the highest average molecular speed?

   (c) A container was with equal amounts H_2O and CO_2 and the pressure was measured. Assuming ideal behavior, if SF_6 was then added such that the final pressure was four times that before its addition, the final number of moles of SF_6 is how many times that of H_2O?

   (d) If a container held equal molar amounts of all three greenhouse gases, but was then compressed decreasing the volume within, which of the gases has the highest, final partial pressure assuming ideal behavior?

   (e) Gas molecules exert pressure through collisions with the walls of the containers they are in. If a container is filled with equal amounts of each of the greenhouse gases, which gas undergoes the fewest collisions with the wall of the container per unit time?

CHAPTER 6 ANSWERS AND EXPLANATIONS

MULTIPLE-CHOICE QUESTIONS

1. **D** Nitrogen gas makes up about 78 percent of the gas in the earth's atmosphere.

2. **E** Density is a measure of grams per liter. CO_2 has the greatest molecular mass (44), so 1 mole of CO_2 will have the most mass in 1 liter and, therefore, the greatest density.

3. **A** According to Graham's law, the lighter the gas, the greater the rate of effusion. H_2 is the lightest gas (MW = 2), so it will have the greatest rate of effusion.

4. **E** CO_2 is the only gas listed that has bonds between atoms of differing electronegativity. The carbon–oxygen bonds in CO_2 are polar, although the linear geometry of the molecule makes the molecule nonpolar overall.

5. **D** N_2 is the only gas listed whose atoms are held together in a triple bond.

6. **C** Remember to convert Celsius to Kelvin (°C + 273 = K).

   27°C = 300 K and 77°C = 350 K.

   From the relationship $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ we get

   $$\frac{(1{,}200 \text{ mmHg})}{(300 \text{ K})} = \frac{P_2}{(350 \text{ K})}$$

   So $P_2 = 1{,}400$ mmHg

7. **B** 8.0 grams of oxygen and 7.0 grams of nitrogen both equal 0.25 moles. Avogadro's law states equal volumes contain equal moles. This makes (A) and (C) wrong.

   (D) This is reversed: the density of oxygen gas would be greater than the density of nitrogen.

   (E) From Graham's law, if the kinetic energies of the two gases are the same, then the molecules of the less massive gas must be moving faster, on average.

8. **D** If the gases are at STP, then the total pressure must be 1.0 atmosphere.

   If $\frac{4}{5}$ of the gas in the sample is nitrogen, then from Dalton's law, $\frac{4}{5}$ of the pressure must be due to the nitrogen.

   So the partial pressure due to nitrogen is $\frac{4}{5}$ (1.0 atm) = 0.8 atm.

9. **E** From Dalton's law, the partial pressure of a gas depends on the number of moles of the gas that are present.

   The total number of moles of gas present is

   1.5 + 3.0 + 0.5 = 5.0 total moles

   If there are 3 moles of nitrogen, then $\frac{3}{5}$ of the pressure must be due to nitrogen.

   $(\frac{3}{5})$(700 mmHg) = 420 mmHg

10. **C** From Dalton's law, the partial pressure of a gas depends on the number of moles of the gas that are present. If the mixture has twice as many moles of helium as neon, then the mixture must be $\frac{1}{3}$ neon. So $\frac{1}{3}$ of the pressure must be due to neon.

$(\frac{1}{3})(1.2 \text{ atm}) = 0.4 \text{ atm}.$

11. **E** From Dalton's law, the partial pressures of nitrogen and water vapor must add up to the total pressure in the container. The partial pressure of water vapor in a closed container will be equal to the vapor pressure of water, so the partial pressure of nitrogen is

781 mmHg – 23 mmHg = 758 mmHg

12. **C** From the gas laws, we know that with constant temperature

$P_1V_1 = P_2V_2$

Solving for P_2, we get

$$P_2 = \frac{P_1V_1}{V_2} = \frac{(6.0 \text{ atm})(24 \text{ L})}{(36 \text{ L})} = 4.0 \text{ atm}$$

13. **C** From kinetic molecular theory, we know that kinetic energy is directly proportional to temperature (think of the expression: $KE = \frac{3}{2}kT$), so (I) is true.

From the gas laws, we know that at constant volume, an increase in temperature will bring about an increase in pressure (think of $\frac{P_1}{T_1} = \frac{P_2}{T_2}$), so (II) is true.

The density of a gas is equal to mass per unit volume, which is not changed by changing temperature, so (III) is not true.

14. **A** Use the relationship

$$\text{Moles} = \frac{\text{liters}}{22.4 \text{ L/mol}}$$

$$\text{Moles of unknown gas} = \frac{11.2 \text{ L}}{22.4 \text{ L/mol}} = 0.500 \text{ moles}$$

$$\text{MW} = \frac{\text{grams}}{\text{mole}}$$

$$\text{MW of unknown gas} = \frac{22.0 \text{ g}}{0.500 \text{ mole}} = 44.0 \text{ grams/mole}$$

That's the molecular weight of CO_2.

15. **E** According to Graham's law, at a given temperature, heavier molecules will have lower average speeds and lighter molecules will have higher speeds.

Helium is the lightest of the three molecules (MW = 4), oxygen is next (MW = 32), and carbon dioxide is the heaviest (MW = 44).

16. **A** The ideal gas laws fail under conditions where gas molecules are packed too tightly together. This can happen as a result of high pressure (I), low temperature (not listed as a choice), or small volume (not listed as a choice).

17. **D** Density is measured in grams per liter. One mole of helium gas has a mass of 4 grams and occupies a volume of 22.4 liters at STP, so the density of helium gas at STP is

$$\frac{4}{22.4}\ \text{g/L}.$$

18. **E** From the ideal gas equation, we know that $PV = nRT$.

Remember to convert 37°C to 310 K.

Then we just rearrange the equation to solve for n.

$$n = \frac{PV}{RT} = \frac{(2.0)(5.0)}{(0.08)(310)}\ \text{moles}$$

19. **D** Remember Graham's law, which relates the average speeds of different gases (and thus their rates of effusion) to their molecular weights.

$$\frac{v_1}{v_2} = \sqrt{\frac{MW_2}{MW_1}}$$

To get a velocity and rate of effusion that is one-third as large as H_2's, we need a molecule with a molecular weight that is nine times as large. The molecular weight of H_2 is 2, so we need a molecule with a molecular weight of 18. That's H_2O.

20. **D** As long as the gases act ideally, the total pressure in the container will only be a function of the number of gas moles present so long as volume and temperature are held constant. The relationship between moles and pressure before and after the reaction, from the ideal gas equation ($PV = nRT$), will be $P_i/n_i = P_f/n_f$ where n_i, P_f, n_f are the initial moles, the final pressure, and the final moles respectively. We can rearrange this equation to $P_f = P_i(n_f/n_i)$, and since there are 2 moles of gas in the products for every 3 moles in the reactants, we can say that $P_f = (2/3)P_i$.

21. **B** From Dalton's law, the partial pressures of nitrogen and water vapor must add up to the total pressure in the container. The partial pressure of water vapor in a closed container will be equal to its vapor pressure, so the partial pressure of nitrogen is

796 mmHg – 55 mmHg = 741 mmHg

22. **E** According to the ideal gas laws, at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles.

We increased the number of grams by a factor of 5 ([0.16][5] = [0.80]). That's the same as increasing the number of moles by a factor of 5. So we must have increased the volume by a factor of 5. Therefore, (5)(1 L) = 5 L.

23. **C** From the ideal gas laws, for a gas sample at constant pressure

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Solving for V_2 we get $V_2 = V_1\dfrac{T_2}{T_1}$

So V_1 is multiplied by a factor of $\dfrac{T_2}{T_1}$

Remember to convert Celsius to Kelvin, $\dfrac{127\,°C + 273}{27\,°C + 273} = \dfrac{400\ K}{300\ K} = \dfrac{4}{3}$

24. **A** We can find the number of moles of gas from $PV = nRT$.

Remember to convert 26°C to 299 K.

Then solve for n.

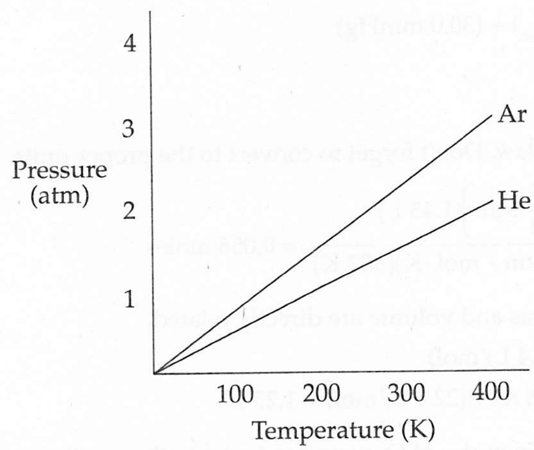
$$n = \frac{PV}{RT} = \frac{(2.0)(6.0)}{(0.08)(299)} \text{ mol}$$

Now, remember

$$\text{MW} = \frac{\text{grams}}{\text{moles}} = \frac{(10 \text{ g})}{\left(\frac{(2.0)(6.0)}{(0.08)(299)} \text{mol}\right)} = \frac{(10)(0.08)(299)}{(2.0)(6.0)} \text{ g/mol}$$

25. **E** All three assumptions are included in kinetic molecular theory.

PROBLEMS



1. (a) Read the graph, and add the two pressures.

$P_{\text{Total}} = P_{\text{He}} + P_{\text{Ar}}$

$P_{\text{Total}} = (1 \text{ atm}) + (1.5 \text{ atm}) = 2.5 \text{ atm}$

(b) Read the pressure (1 atm) at 200 K, and use the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{(1.0 \text{ atm})(2.0 \text{ L})}{(0.082 \text{ L-atm/mol-K})(200 \text{ K})} = 0.12 \text{ moles}$$

(c) Use the definition of a mole.

Molecules = (moles)(6.02×10^{23})

Molecules (atoms) of helium = (0.12)(6.02×10^{23}) = 7.2×10^{22}

(d) Use Graham's law.

$$\frac{v_1}{v_2} = \sqrt{\frac{\text{MW}_2}{\text{MW}_1}}$$

$$\frac{v_{He}}{v_{Ar}} = \sqrt{\frac{MW_{Ar}}{MW_{He}}} = \sqrt{\frac{(40)}{(4)}} = 3.2 \text{ to } 1$$

(e) Use the relationship

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

With T constant

$P_1V_1 = P_2V_2$

(1.5 atm)(2.0 L) = P_2(1.0 L)

P_2 = 3.0 atm

2. (a) Use Dalton's law.

$P_{Total} = P_{Oxygen} + P_{Water}$

(755 mmHg) = (P_{Oxygen}) + (30.0 mmHg)

P_{Oxygen} = 725 mmHg

(b) Use the ideal gas law. Don't forget to convert to the proper units.

$$n = \frac{PV}{RT} = \frac{\left(\frac{725}{760}\text{ atm}\right)(1.45\text{ L})}{(0.082\text{ L-atm / mol-K})(302\text{ K})} = 0.056 \text{ moles}$$

(c) At STP, moles of gas and volume are directly related.

Volume = (moles)(22.4 L/mol)

Volume of O_2 = (0.056 mol)(22.4 L/mol) = 1.25 L

(d) We know that 0.056 moles of O_2 were produced in the reaction.

From the balanced equation, we know that for every 3 moles of O_2 produced, 2 moles of $KClO_3$ are consumed. So there are $\frac{2}{3}$ as many moles of $KClO_3$ as O_2.

Moles of $KClO_3$ = $(\frac{2}{3})$(moles of O_2)

Moles of $KClO_3$ = $(\frac{2}{3})$(0.056 mol) = 0.037 moles

Grams = (moles)(MW)

Grams of $KClO_3$ = (0.037 mol)(122 g/mol) = 4.51 g

Essays

3. (a) The partial pressures depend on the number of moles of gas present. Because the number of moles of the two gases are the same, the partial pressures are the same.

(b) O_2 has the greater density. Density is mass per unit volume. Both gases have the same number of moles in the same volume, but oxygen has heavier molecules, so it has greater density.

(c) Concentration is moles per volume. Both gases have the same number of moles in the same volume, so their concentrations are the same.

(d) According to kinetic-molecular theory, the average kinetic energy of a gas depends only on the temperature. Both gases are at the same temperature, so they have the same average kinetic energy.

(e) H_2O will deviate most from ideal behavior. Ideal behavior for gas molecules assumes that there will be no intermolecular interactions.

H_2O is polar, and O_2 is not. H_2O undergoes hydrogen bonding while O_2 does not. So H_2O has stronger intermolecular interactions, which will cause it to deviate more from ideal behavior.

(f) H_2O has the greater average molecular speed. From Graham's law, if two gases are at the same temperature, the one with the smaller molecular weight (MW of $H_2O = 18$, MW of $O_2 = 32$) will have the greater average molecular speed.

4. (a) H_2O deviates most drastically from ideal behavior due to very strong hydrogen bonding interactions. A truly ideal gas has no interactions with surrounding molecules. These hydrogen bonding interactions are why H_2O is a liquid at room temperature and the others are gases.

(b) As per Graham's law, the species with the lowest molecular weight would have the greatest average molecular speed. H_2O, MW = 18 g/mol, is the lightest, and hence would be the fastest.

(c) By Dalton's law of partial pressures, if SF_6 was added until the pressure was 4 times the pressure with just (an equal mixture of) H_2O and CO_2, then there must be 3 times as much SF_6 as H_2O and CO_2 combined. This means that, in the end, there 6 times more SF_6 than H_2O. For example, imagine starting with 1 atm each of CO_2 and H_2O. In this case the initial pressure is 2 atm, and to get to 8 atm one must add 6 atm of SF_6 which is 6 times that of the H_2O present.

(d) All three gases will have identical partial pressures. The equation for partial pressure of any component a is $P_a = (P_{total})(X_a)$. Since we know that each has the same mol fraction the total pressure doesn't matter, and each will have the same partial pressure, relative to one another.

(e) SF_6 will have the fewest collisions per unit time. Since the kinetic theory assumes that the gas molecules are moving randomly, the one with the most collisions with the walls of the containers will be the one moving the fastest. On the other hand, the slowest molecule will have the fewest collisions. The slowest molecule will be the most massive one, as per Graham's law, which in this case is SF_6.