Name ____________________ Section ________ Date ________

Chapter 13: Basic Review Worksheet

1. What are some of the general properties of gases that distinguish them from liquids and solids?

2. What is the SI unit of pressure? What units of pressure are commonly used in the United States? Why are these common units more convenient to use than the SI unit?

3. Convert 1.20 atm to units of mm Hg, torr, and pascals.

4. Your textbook gives several definitions and formulas for Boyle's law for gases. Write, in your *own* words, what this law really tells us about gases.

5. What does Charles's law tell us about how the volume of a gas sample varies as the temperature of the sample is changed?

6. What temperature scale is defined with its lowest point as the absolute zero of temperature? What is absolute zero in Celsius degrees?

7. What does Avogadro's law tell us about the relationship between the volume of a sample of gas and the number of molecules the gas contains?

8. What do we mean specifically by an *ideal* gas?

9. What is the numerical value and what are the specific units of the universal gas constant, R? Why is close attention to *units* especially important when doing ideal gas law calculations?

10. A sample of gas in a 10.0-L container exerts a pressure of 565 mm Hg. Calculate the pressure exerted by the gas if the volume is changed to 15.0 L at constant temperature.

11. A sample of gas in a 5.00-L container at 35.0°C is heated at constant pressure to a temperature of 70.0°C at constant pressure. Determine the volume of the heated gas.

12. A 4.50 mol sample of a gas occupies a volume of 34.6 L at a particular temperature and pressure. What volume does 2.50 mol of the gas occupy at these same conditions of pressure and temperature?

13. A sample of gas at 24°C occupies a volume of 3.45 L and exerts a pressure of 2.10 atm. The gas is cooled to −12°C and the pressure is increased to 5.20 atm. Determine the new volume occupied by the gas.

14. What mass of helium gas exerts a pressure of 1.20 atm in a volume of 5.40 L at a temperature of 27°C?

15. Dalton's law of partial pressures concerns the properties of mixtures of gases. What is meant by the *partial pressure* of an individual gas in a mixture?

16. A 2.50 g sample of neon gas is added to a 5.00 g sample of argon gas in a 10.0-L container at 23°C. Calculate the partial pressure of each gas and the total pressure of the mixture.

17. A sample of oxygen gas is collected over water at 27°C. The total pressure is 0.95 atm and the water vapor pressure at 27°C is 26.7 torr. Determine the partial pressure of the oxygen gas collected.

18. When calcium carbonate is heated strongly, carbon dioxide gas is evolved:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Determine the volume occupied by the carbon dioxide produced by the decomposition of 23.5 g of calcium carbonate. The carbon dioxide is collected at 1.10 atm and 24°C.

19. What does "STP" stand for? What conditions correspond to STP?

20. Under what conditions of pressure and temperature does a gas behave most ideally? Explain.

Name ____________________________ Section __________ Date __________

Chapter 13: Review Worksheet

1. Describe a *manometer* and explain how such a device can be used to measure the pressure of gas samples.

2. Write a mathematical expression for Boyle's law and explain it. Sketch the general shape of a graph of pressure versus volume for an ideal gas and explain it.

3. When using Boyle's law in solving problems in the textbook, you may have noticed that questions were often qualified by stating "the temperature and amount of gas remain the same." Why is this qualification necessary?

4. How does the volume-temperature relationship of Charles's law differ from the volume-pressure relationship of Boyle's law?

5. For Charles's law to hold true, why must the pressure and amount of gas remain the same?

6. Does Avogadro's law describe a direct or an inverse relationship between the volume and the number of moles of gas?

7. Why must the temperature and pressure be held constant for valid comparison using Avogadro's law?

8. Explain how the *ideal gas law* is actually a combination of Boyle's, Charles's, and Avogadro's gas laws.

9. A sample of gas in a 25.0-L container exerts a pressure of 3.20 atm. Calculate the pressure exerted by the gas if the volume is changed to 45.0-L at constant temperature.

10. A sample of gas in a 21.5-L container at 45°C is cooled at constant pressure to a temperature of -37°C at constant pressure. Determine the volume of the cooled gas.

11. A 32.8 g sample of hydrogen gas occupies a volume of 21.6 L at a particular temperature and pressure. What volume does 12.3 g of hydrogen gas occupy at the same pressure and temperature?

12. A sample of gas at 38°C occupies a volume of 2.97 L and exerts a pressure of 3.14 atm. The gas is heated to 118°C and the volume is decreased to 1.04 L. Determine the new pressure exerted by the gas.

13. What mass of oxygen gas exerts a pressure of 475 mm Hg in a volume of 1.25 L at a temperature of -22°C?

14. How does the *total pressure* of a gaseous mixture depend on the partial pressures of the individual gases in the mixture?

15. Many common laboratory preparations of gaseous substances involve collecting the gas produced by displacement of water from a receiving container. How is Dalton's law of partial pressures used in determining the partial pressure of the prepared gas in such an experiment?

16. A 12.5 g sample of oxygen gas is added to a 25.0 g sample of nitrogen gas in a 25.0-L container at 28°C. Calculate the partial pressure of each gas and the total pressure of the mixture.

17. A sample containing 0.80 mol of oxygen gas is collected over water at 30.0°C. The total pressure is 1.10 atm and the water vapor pressure at 30.0°C is 31.8 torr. Determine the volume of the oxygen gas.

18. Without consulting your textbook, list and explain the main postulates of the kinetic molecular theory for gases. How do these postulates help us account for the following properties of a gas: the pressure of the gas and why the pressure of the gas increases with increased temperature; the fact that a gas fills its entire container, and the fact that the volume of a given sample of gas increases as its temperature increased.

19. Zinc metal reacts with hydrochloric acid according to the following unbalanced equation:

$$Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

A 10.0-g sample of zinc is reacted with 0.200 mol of HCl. Determine the volume occupied by the hydrogen gas collected at 755 mm Hg and 22°C.

20. Do gases behave most ideally at STP? Explain.

Name ______________________ Section ________ Date ________

Chapter 13: Challenge Review Worksheet

1. One of the most obvious properties of gaseous materials is the pressure they exert on their surroundings. In particular, the pressure exerted by the atmospheric gases is important. How does the pressure of the atmosphere arise, and how is this pressure commonly measured?

2. Explain how the concept of absolute zero came about through Charles's studies of gases. **Hint:** What would happen to the volume of a gas sample at absolute zero (if the gas did not liquefy first)?

3. A 12.4-g sample of helium gas occupies a volume of 23.5 L at a certain temperature and pressure. What volume does a 56.2 g sample of neon gas occupy at these conditions of temperature and pressure?

4. If the volume of a given amount of gas is doubled and the Celsius temperature is doubled, what will happen to the pressure? Explain using the gas laws, and the kinetic molecular theory.

5. How does Dalton's law help us realize that for an ideal gas sample, the volume of an individual molecule is insignificant compared with the bulk volume of the sample?

6. Zinc and magnesium metal each react with hydrochloric acid according to the following equation (where M represents either Zn or Mg):

$$M(s) + 2HCl(aq) \rightarrow MCl_2(aq) + H_2(g)$$

A 10.00-g mixture of zinc and magnesium produces 5.74 L of hydrogen gas at 1.10 atm and 27°C. Determine the percent magnesium in the original mixture.

7. Why aren't gases ideal?

In each of these structures, the central atom (C, N or O) is surrounded by four pairs (an octet) of electrons. According to the VSEPR theory, the four pairs of electrons repel each other and orient themselves in space as far away from each other as possible. This leads to a tetrahedron orientation of the electron pairs, separated by angles of 109.5°. For CH_4, each of the four pairs of electrons around the C atom is a bonding pair. We therefore say that CH_4 itself has a tetrahedral geometry, with H-C-H bond angles of 109.5°. For NH_3, however, although there are four tetrahedrally arranged electron pairs around the nitrogen atom, only three of these pairs are bonding pairs: there is a lone pair on the nitrogen atom. We describe the geometry of NH_3 as a trigonal pyramid: the three hydrogen atoms lie below the nitrogen atom in space as a result of the presence of the lone pair on nitrogen. The H-N-H bond angles are slightly less than the tetrahedral angle of 109.5°. Finally, only two of these pairs are bonding pairs: there are two lone electron pairs on the oxygen atom. We describe the geometry of H_2O as bent, v-shaped, or nonlinear: the presence of the lone pairs makes the H-O-H bond angle not 180° (linear) but somewhat less than the tetrahedral angle of 109.5°.

7.

Number of valence pairs	Bond angle	Example(s)
2	180°	BeF_2, BeH_2
3	120°	BCl_3
4	109.5°	CH_4, CCl_4, GeF_4

8. In predicting the geometric structure of a molecule, we treat a double (or triple) bond as a single entity (as if it were a single pair of electrons). This approach is reasonable, since all the bonding electrons between two atoms must be present in the same region of space between the atoms (whether one, two, or three electron pairs). For example, if we write the Lewis structure of acetylene, C_2H_2

   H: C:::C:H

   We realize that each carbon atom in the molecule has effectively only two "things" (termed "repulsive units" in the text) attached to it: a bonding pair of electrons (which bonds the H atom) and a triple bond (which bonds the other C atom). Since there are effectively only two things (repulsive units) attached to each carbon atom, we would expect the bond angles for each carbon atom to be 180°, which makes the molecule linear.

Chapter 13: Basic Review Worksheet

1. Gases have no fixed volume or shape, but take on the shape and volume of the container in which they are confined. This is in contrast to solids and liquids: A sample of solid has its own intrinsic volume and shape, and is very incompressible. A sample of liquid has an intrinsic volume, but does not take on the shape of its container.

2. The SI unit of pressure is the pascal, but this unit is almost never used in everyday situations because it is too small to be practical. Rather, we tend to use units of pressure that are based on the simple instruments used to measure pressures, the mercury barometer and manometer. The mercury barometer, used for measuring the pressure of the atmosphere, consists of a column of mercury that is held in a vertical glass tube by the atmosphere. The pressure of the atmosphere is indicated by the height of the mercury in the long tube (relative to the surface of the mercury in the reservoir). As the atmospheric pressure changes, the height of the mercury column changes. The height of the mercury column is given in radio and TV weather reports in inches in mercury,

but most scientific applications would quote the height in millimeters of mercury (mm Hg, torr). Pressures are also quoted in standard atmospheres, where 1 atm is equivalent to a pressure of 760 mm Hg.

3. $1.20\text{ atm} \times \dfrac{760\text{ mm Hg}}{1\text{ atm}} = 912\text{ mm Hg}$

   $1.20\text{ atm} \times \dfrac{760\text{ torr}}{1\text{ atm}} = 912\text{ torr}$

   $1.20\text{ atm} \times \dfrac{101{,}325\text{ Pa}}{1\text{ atm}} = 122{,}000\text{ Pa}$

4. Answers will vary, but consider the following: Boyle's law basically says that the volume of a gas sample will decrease if you squeeze harder on it. Imagine squeezing hard on a tennis ball with your hand: the ball collapses as the gas inside it is forced into a smaller volume by your hand. The students should remember that the temperature and amount of gas (moles) must remain the same for Boyle's law to hold true.

5. Charles's law basically says that if you heat a sample of gas, the volume of the sample will increase. That is, when the temperature of a gas is increased, the volume of the gas also increases (assuming the pressure and amount of gas remain the same).

6. The Kelvin or absolute temperature scale is defined with absolute zero as its lowest temperature, with all temperature positive relative to this point. The size of the Kelvin degree was chosen to be the same size as the Celsius degree. Absolute zero (0 K) corresponds to –273°C.

7. Avogadro's law tells us that, with all other things being equal, two moles of gas have twice the volume of one mole of gas. That is, the volume of a sample of gas is directly proportional to the number of moles or molecules of gas present (at constant temperature and pressure).

8. An ideal gas is defined to be a gas that obeys the ideal gas law.

9.

$$R = \frac{0.08206\text{ L atm}}{\text{mol K}}$$

   Although it is always important to pay attention to the units when solving a problem, this is especially important when solving gas problems involving the universal gas constant, R. The numerical value of 0.08206 for R applies only when the properties of the gas sample are given in the units specified for the constant: the volume in liters (not mL,) the pressure in atmospheres (not mm Hg, torr, or Pa), the amount of gas moles (not g), and the temperature in units of Kelvin (not °F or °C).

10. Use $P_1V_1 = P_2V_2$
    (565 mmHg) (10.0 L) = P_2 (15.0 L)
    377 mm Hg = P_2

11. Use $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$\frac{5.00\text{ L}}{[(35+273)\text{K}]} = \frac{V_2}{[(70+273)\text{K}]}$$

$5.57\text{ L} = V_2$

12. Use $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

$$\frac{34.6\text{ L}}{4.50\text{ mol}} = \frac{V_2}{2.50\text{ mol}}$$

$19.2\text{ L} = V_2$

13. Use $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$\frac{(2.10\text{ atm})(3.45\text{ L})}{[(24+273)\text{K}]} = \frac{(5.20\text{ atm})(V_2)}{[(-12+273)\text{K}]}$$

$1.22\text{ L} = V_2$

14. Use $PV = nRT$, $n = \frac{PV}{RT}$

$$n = \frac{(1.20\text{ atm})(5.40\text{ L})}{(0.08206\frac{\text{L atm}}{\text{mol K}})[(27+273)\text{K}]} = 0.263\text{ mol He}$$

$$0.263\text{ mol He} \times \frac{4.003\text{g He}}{1\text{ mol He}} = 1.05\text{g He}$$

15. The "partial" pressure of an individual gas in a mixture of gases represents the pressure the gas would have in the same container at the same temperature if it were the only gas present.

16. $2.50\text{g Ne} \times \frac{1\text{ mol Ne}}{20.18\text{g Ne}} = 0.124\text{ mol Ne}$

$$5.00\text{g Ar} \times \frac{1\text{ mol Ar}}{39.95\text{g Ar}} = 0.125\text{ mol Ar}$$

$$P_{Ne} = \frac{(0.124\,\text{mol})(0.08206\,\frac{\text{L atm}}{\text{mol K}})[(23+273)\text{K}]}{(10.0\,\text{L})} = 0.301\text{ atm}$$

$$P_{Ar} = \frac{(0.125\,\text{mol})(0.08206\,\frac{\text{L atm}}{\text{mol K}})[(23+273)\text{K}]}{(10.0\,\text{L})} = 0.304\text{ atm}$$

$$P_{TOTAL} = 0.605\text{ atm}$$

17. $0.95\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 722\text{ torr}$

$$P_{TOTAL} = P_{O_2} + P_{H_2O}$$

$$722\text{ torr} = P_{O_2} + 26.7\text{ torr}$$

$$P_{O_2} = 695\text{ torr}$$

18. $23.5\text{g } CaCO_3 \times \frac{1\text{ mol } CaCO_3}{100.09\text{g } CaCO_3} \times \frac{1\text{ mol } CO_2}{1\text{ mol } CaCO_3} = 0.235\text{ mol } CO_2$

$$V = \frac{(0.235\,\text{mol } CO_2)(0.08206\,\frac{\text{L atm}}{\text{mol K}})[(24+273)\text{K}]}{(1.10\,\text{atm})} = 5.21\text{ L}$$

19. The abbreviation "STP" stands for "Standard Temperature and Pressure". STP corresponds to a temperature of 0°C and a pressure of 1 atm. These conditions were chosen for comparisons of gas samples because they are easy to reproduce in any laboratory (an equilibrium mixture of ice and water has a temperature of 0°C, and the pressure in most laboratories is very near to 1 atm).

20. Gases behave most ideally at conditions of low pressure and high temperature. Two of the premises of the kinetic molecular theory are that gas particles exert no forces on each other and the volume of gas particles is negligible compared to the volume of the container. These conditions are more closely approached as the pressure of the gas is decreased (less gas per volume) and the temperature of the gas is increased (greater average kinetic energy results in faster moving particles).

Chapter 13: Review Worksheet

1. While the barometer is used to measure atmospheric pressure, a device called a mercury manometer is used to measure the pressure of samples of gas in the laboratory. A manometer consists basically of a U-shaped tube filled with mercury, with one arm of the U open to the atmosphere. If the pressure of the gas sample equals atmospheric pressure then the mercury levels will be the same in both sides of the U. If the pressure of the gas is not the same as the

atmospheric pressure, then the difference in height of the mercury levels can be sued to determine by how many mm Hg the pressure of the gas sample differs from atmospheric pressure.

2. The expression P x V = constant is Boyle's law. In order for the product (P × V) to remain constant, if one of these terms increases the other must decrease. A second formulation of Boyle's law is one more commonly used in solving problems:

$$P_1 \times V_1 = P_2 \times V_2$$

With this second formulation, we can determine pressure-volume information about a given sample under two sets of conditions. These two mathematical formulas are just two different ways of saying the same thing: if the pressure on a sample of gas is increased, the volume of the sample of gas will decrease. A graph of Boyle's law data is given as Figure 13.5: this sort of graph ($xy = k$) is known to mathematicians as a hyperbola.

3. The qualification is necessary because the volume of a gas sample is dependent on all its properties. The properties of a gas are all inter-related (as shown by the ideal gas law, $PV = nRT$). If we want to use one of the derivative gas laws (Boyle's, Charles's, or Avogadro's gas laws), which isolate how the volume of a gas sample varies with just one of its properties, then we must keep all the other properties constant while that one property is studied.

4. Charles's law is a direct proportionality when the temperature is expressed in Kelvins (if you increase T, this increases V), whereas Boyle's law is an inverse proportionality (if you increase P, this decreases V).

5. Charles's law only holds true if the amount of gas remains the same (for example, the volume of a gas sample would increase if there were more gas present) and also if the pressure remains the same (a change in pressure also changes the volume of a gas sample).

6. Avogadro's law is a direct proportionality: the greater the number of gas molecules in a sample, the larger the sample's volume will be.

7. Comparing the volumes of two samples of the same gas to determine the relative of the amount of gas present in the samples requires that the two samples of gas are at the same pressure and temperature: the volume of a sample of gas would vary with either temperature (according to Charles's law) or pressure (according to Boyle's law), or both. Avogadro's law holds true for comparing gas samples that are under the same conditions of pressure and temperature.

8. Boyle's law states that the volume of a gas is inversely proportional to its pressure (at constant temperature for a fixed amount of gas):

$$V = (\text{constant})/P$$

Charles's law indicates that the volume of a gas sample is related to its temperature (at constant pressure for a fixed amount of gas):

$$V = (\text{constant}) \times T$$

Avogadro's law states that the volume of a gas sample is proportional to the number of moles of gas (at constant pressure and temperature):

$$V = (\text{constant}) \times n$$

We can combine all these relationships (and constants) to show how the volume of a gas is proportional to all its properties simultaneously:

$$V = (\text{constant}) \times \frac{T \times n}{P}$$

This can be arranged to the familiar form of the ideal gas law: PV = nRT.

9. Use $P_1V_1 = P_2V_2$
$(3.2\text{ atm})(25.0\text{ L}) = P_2(45.0\text{ L})$
$1.78\text{ atm} = P_2$

10. Use $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$\frac{21.5\text{ L}}{[(45+273)\text{K}]} = \frac{V_2}{[(-37+273)\text{K}]}$$

$16.0\text{ L} = V_2$

11. Since both samples are hydrogen, the number of moles is directly related to mass, thus use

$$\frac{V_1}{\text{mass}_1} = \frac{V_2}{\text{mass}_2}$$

$$\frac{21.6\text{ L}}{32.8\text{g}} = \frac{V_2}{12.3\text{g}}$$

$8.10\text{ L} = V_2$

12. Use $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$\frac{(3.14\text{ atm})(2.97\text{ L})}{[(38+273)\text{K}]} = \frac{(P_2)(1.04\text{ L})}{[(118+273)\text{K}]}$$

$11.3\text{ atm} = P_2$

13. Use PV = nRT, $n = \dfrac{PV}{RT}$

$$P = 475\text{mm Hg} \times \frac{1\text{ atm}}{760\text{ mm Hg}} = 0.625\text{ atm}$$

$$n = \frac{(0.625\text{ atm})(1.25\text{ L})}{(0.08206\frac{\text{L atm}}{\text{mol K}})[(-22+273)\text{K}]} = 0.0379\text{ mol } O_2$$

$$0.0379\text{ mol } O_2 \times \frac{32.00\text{g } O_2}{1\text{ mol } O_2} = 1.21\text{ g } O_2$$

14. The total pressure in a mixture of gases is the sum of the individual partial pressures of the gases present in a mixture.

15. When a gas is collected by displacement of water from a container, the gas becomes saturated with water vapor. The collected gas is actually a mixture of the desired gas and water vapor. To determine the partial pressure of the desired gas in the mixture, it is necessary to subtract the pressure of water vapor from the total pressure of the sample

$$P_{gas} = P_{total} - P_{water\ vapor}$$

Dalton's law of partial pressures states that the total pressure in a mixture of gases is the sum of the partial pressures of the components of the mixture. Since the saturation pressure of water vapor is a function only of temperature, such water vapor pressures are conveniently tabulated (see Table 13.2 in the text).

16. $12.5\text{g } O_2 \times \dfrac{1\text{ mol } O_2}{32.00\text{g } O_2} = 0.391\text{ mol } O_2$

$$25.0\text{g } N_2 \times \frac{1\text{ mol } N_2}{28.02\text{g } N_2} = 0.892\text{ mol } N_2$$

$$P_{O_2} = \frac{(0.391\text{ mol})(0.08206\frac{\text{L atm}}{\text{mol K}})[(28+273)\text{K}]}{(25.0\text{ L})} = 0.386\text{ atm}$$

$$P_{N_2} = \frac{(0.892\text{ mol})(0.08206\frac{\text{K atm}}{\text{mol K}})[(28+273)\text{K}]}{(25.0\text{ L})} = 0.881\text{ atm}$$

$P_{TOTAL} = 1.267$ atm

17. $P_{H_2O} = 31.8 \text{ torr} \times \dfrac{1 \text{ atm}}{760 \text{ torr}} = 0.0418 \text{ atm}$

$P_{O_2} = 1.10 \text{ atm} - 0.0418 \text{ atm} = 1.06 \text{ atm}$

$$V = \frac{(0.80 \text{ mol})(0.08206 \frac{\text{K atm}}{\text{mol K}})[(30 + 273)\text{K}]}{(1.06 \text{ atm})}$$

V = 18.8 L

18. The main postulates of the kinetic-molecular theory for gases are as follows: (a) gases consist of tiny particles (atoms or molecules), and the size of these particles themselves is negligible compared to the bulk volume of a gas sample; (b) the particles in a gas are in constant random motion, colliding with the walls of the container; (c) the particles in a gas sample do not exert any attractive or repulsive forces on one another; (d) the average kinetic energy of the particles in a sample of gas is directly related to the absolute temperature of the gas sample. The pressure exerted by a gas is a result of the molecules colliding with (and pushing on) the walls of the container. The pressure increases with temperature because at a higher temperature, the molecules are moving faster and hit the walls of the container with greater force. A gas fills whatever volume is available to it because the molecules in a gas are in constant random motion: if the motion of the molecules is random, they eventually will move out into whatever volume is available until the distribution of molecules is uniform. At constant pressure, the volume of a gas sample increases as the temperature is increased because with each collision having greater force, the container must expand so that the molecules hit the walls less frequently to maintain the same pressure.

19. The balanced equation is $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$$10.0 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38\text{g Zn}} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Zn}} = 0.153 \text{ mol } H_2$$

$$0.200 \text{ mol HCl} \times \frac{1 \text{ mol } H_2}{2 \text{ mol HCl}} = 0.100 \text{ mol } H_2$$

Thus, 0.100 mol H_2 is produced since HCl is limiting.

$$V = \frac{(0.100 \text{ mol})(0.08206 \frac{\text{K atm}}{\text{mol K}})[(22 + 273)\text{K}]}{(755 \text{ atm} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}})}$$

V = 2.44 L

20. Gases do not behave most ideally at STP. Standard temperature and pressure (0°C, 1 atm) is chosen as standard because it is easy to produce these conditions in the lab, not because gases behave most ideally at these conditions. Many students believe this to be the case, however.

Chapter 13: Challenge Review Worksheet

1. The pressure exerted by the atmosphere is due to the several-mile-thick layer of gases above the surface of the earth pressing down on us. Atmospheric pressure has traditionally been measured with a mercury barometer. A mercury barometer usually consists of a glass tube that is sealed at one end and filled with mercury. The tube is then inverted into an open reservoir also containing mercury. When the tube is inverted, most of the mercury does not fall out of the tube. Since the reservoir of mercury is open to the atmosphere, the atmospheric pressure on the surface of the mercury in the reservoir is enough pressure to hold the bulk of the mercury in the glass tube. The pressure of the atmosphere is sufficient, on average, to support a column of mercury 76 cm (760 mm) high in the tube.

2. The volume of a gas sample changes by the same factor (i.e., linearly) for each Celsius degree its temperature is changed (for a fixed amount of gas at a constant pressure). Charles realized that, if a gas were cooled, the volume of a gas sample would decrease by a constant factor for each degree the temperature was lowered. When Charles plotted his experimental data, and extrapolated the linear data to very low temperatures that he could not measure experimentally, he realized that there would be an ultimate temperature no matter what gas sample was used for the experiment. This temperature—where the volume of an ideal gas sample would approach zero as a limit—is called the absolute zero of temperature. Unlike the Fahrenheit and Celsius temperature scales, which were defined by humans with experimentally convenient reference points, the absolute zero of temperature is a fundamental, natural reference point for the measurement of temperatures.

3. $12.4\text{g He} \times \dfrac{1\text{ mol He}}{4.003\text{g He}} = 3.10\text{ mol He}$

   $56.2\text{g Ne} \times \dfrac{1\text{ mol Ne}}{20.18\text{g Ne}} = 2.78\text{ mol Ne}$

   use $\dfrac{V_1}{n_1} = \dfrac{V_2}{n_2}$ (the identity of the gas does not matter)

   $$\frac{23.5\text{ L}}{3.10\text{ mol}} = \frac{V_2}{2.78\text{ mol}}$$

   $21.1\text{ L} = V_2$

4. Doubling the volume should cut the pressure in half (Boyle's law) and doubling the Celsius temperature will increase the pressure but not double it (the Kelvin temperature must be doubled to double the pressure). Thus, doubling the volume and Celsius temperature will cause a decrease in pressure. Boyle's law holds true because at twice the volume the gas particles will

make half the number of impacts with the walls, thus the pressure is halved. As the temperature is increased, the average kinetic energy of the particles is increased (though not doubled since average kinetic energy is related to the Kelvin temperature). Thus, the particles move faster and hit the walls with a greater force and frequency and the pressure is increased.

5. Because the partial pressures of the gases in a mixture are additive (i.e., the total pressure is the sum of the partial pressures), this suggests that the total pressure in a container is a function only of the number of molecules present, and not of the identity of the molecules or any other property of the molecules (such as their inherent atomic size).

6. $$\text{moles } H_2 = \frac{(1.10\text{ atm})(5.74\text{ L})}{(0.08206\frac{\text{L atm}}{\text{mol K}})[(27+273)\text{K}]} = 0.256\text{ mol } H_2$$

mass Zn + mass Mg = 10.0g

mol Zn + mol Mg = 0.256 mol

$$\frac{\text{mass Zn}}{\text{mol Zn}} = 65.38\text{ g/mol}\ ;\ \frac{\text{mass Mg}}{\text{mol Mg}} = 24.31\text{ g/mol}$$

(a) (65.38)(mol Zn) + (24.31)(mol Mg) = 10.00

(b) mol Zn + mol Mg = 0.256

multiply (b) by "24.31" and subtract from (a)

(41.07)(mol Zn) = 3.78

mol Zn = 0.0920 mol

mol Mg = 0.164 mol

$$\text{mass Mg} = 0.164\text{ mol Mg} \times \frac{24.31\text{ g}}{1\text{ mol}} = 3.99\text{ g Mg}$$

$$\frac{3.99\text{ g Mg}}{10.00\text{g}} \times 100\% = 39.9\%\text{ Mg}$$

7. Gases are not ideal because the particles making up a sample of gas exert forces on each other and have a finite volume.