

Chemistry 2e
9: Gases
9.1: Gas Pressure

1. Why are sharp knives more effective than dull knives? (Hint: Think about the definition of pressure.)

Solution

The cutting edge of a knife that has been sharpened has a smaller surface area than a dull knife. Since pressure is force per unit area, a sharp knife will exert a higher pressure with the same amount of force and cut through material more effectively.

3. Why should you roll or belly crawl rather than walk across a thinly frozen pond?

Solution

Lying down distributes your weight over a larger surface area, exerting less pressure on the ice compared to standing up. If you exert less pressure, you are less likely to break through thin ice.

5. A typical barometric pressure in Denver, Colorado, is 615 mm Hg. What is this pressure in atmospheres and kilopascals?

Solution

Convert 615 mm Hg to atmospheres using 760 mm Hg = 1 atm. Use 1 atm = 101.325 kPa in the second part.

$$615 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.809 \text{ atm}$$

$$0.809 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 82.0 \text{ kPa}$$

7. Canadian tire pressure gauges are marked in units of kilopascals. What reading on such a gauge corresponds to 32 psi?

Solution

$$32.0 \text{ lb in}^{-2} \times \frac{1 \text{ atm}}{14.7 \text{ lb in}^{-2}} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 2.2 \times 10^2 \text{ kPa}$$

9. The pressure of the atmosphere on the surface of the planet Venus is about 88.8 atm. Compare that pressure in psi to the normal pressure on earth at sea level in psi.

Solution

Identify: 14.7 psi = 1 atm

$$88.8 \text{ atm} \times \frac{14.7 \text{ psi}}{1 \text{ atm}} = 1.30 \times 10^3 \text{ psi}$$

11. Consider this scenario and answer the following questions: On a mid-August day in the northeastern United States, the following information appeared in the local newspaper: atmospheric pressure at sea level 29.97 in., 1013.9 mbar.

(a) What was the pressure in kPa?

(b) The pressure near the seacoast in the northeastern United States is usually reported near 30.0 in. Hg. During a hurricane, the pressure may fall to near 28.0 in. Hg. Calculate the drop in pressure in torr.

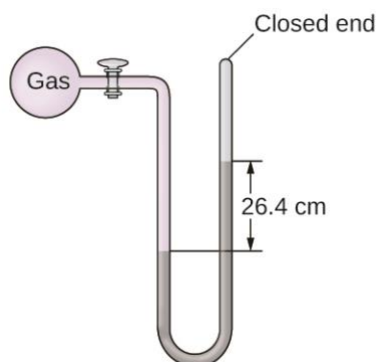
Solution

$$(a) 29.97 \text{ in. Hg} \times \frac{101.325 \text{ kPa}}{29.92 \text{ in. Hg}} = 101.5 \text{ kPa}; (b) 28.0 \text{ in. Hg} \times \frac{760 \text{ torr}}{29.92 \text{ in. Hg}} = 711 \text{ torr};$$

$$762 - 711 = 51 \text{ torr drop}$$

13. The pressure of a sample of gas is measured at sea level with a closed-end manometer. The liquid in the manometer is mercury. Determine the pressure of the gas in:

- (a) torr
(b) Pa
(c) bar



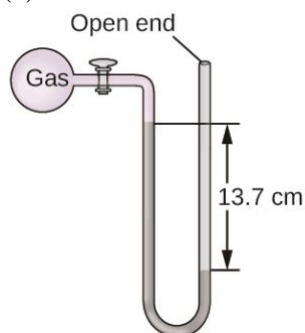
Solution

$$(a) \cancel{26.4 \text{ cm}} \times \frac{10 \cancel{\text{ mm}}}{1 \cancel{\text{ cm}}} \times \frac{1 \text{ torr}}{1 \cancel{\text{ mm}}} = 264 \text{ torr}; (b) 264 \text{ torr} \times \frac{101,325 \text{ Pa}}{760 \text{ torr}} = 35,200 \text{ Pa}; (c)$$

$$264 \cancel{\text{ torr}} \times \frac{1.01325 \text{ bar}}{760 \cancel{\text{ torr}}} = 0.352 \text{ bar}$$

15. The pressure of a sample of gas is measured at sea level with an open-end mercury manometer. Assuming atmospheric pressure is 760.0 mm Hg, determine the pressure of the gas in:

- (a) mm Hg
(b) atm
(c) kPa



Solution

The pressure of the gas equals the hydrostatic pressure due to the pressure of the atmosphere at sea level minus a column of mercury of height 13.7 cm. The pressure on the left is due to the gas and the pressure on the right is due to the atmospheric pressure minus 13.7 cm Hg). (a) In mm Hg, this is: 760 mm Hg – 137 mmHg = 623 mm Hg; (b)

$$623 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.820 \text{ atm}; (c) 0.820 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 83.1 \text{ kPa}$$

17. How would the use of a volatile liquid affect the measurement of a gas using open-ended manometers vs. closed-end manometers?

Solution

With a closed-end manometer, no change would be observed, since the vaporized liquid would contribute equal, opposing pressures in both arms of the manometer tube. However, with an open-ended manometer, a higher pressure reading of the gas would be obtained than expected, since $P_{\text{gas}} = P_{\text{atm}} + P_{\text{vol liquid}}$.

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Chemistry 2e

9: Gases

9.2: Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law

19. Explain how the volume of the bubbles exhausted by a scuba diver (Figure 9.16) change as they rise to the surface, assuming that they remain intact.

Solution

As the bubbles rise, the pressure decreases, so their volume increases as suggested by Boyle's law.

21. An alternate way to state Avogadro's law is "All other things being equal, the number of molecules in a gas is directly proportional to the volume of the gas."

(a) What is the meaning of the term "directly proportional?"

(b) What are the "other things" that must be equal?

Solution

(a) The number of particles in the gas increases as the volume increases. This relationship may be written as $n = \text{constant} \times V$. It is a direct relationship. (b) The temperature and pressure must be kept constant.

23. How would the graph in Figure 9.13 change if the number of moles of gas in the sample used to determine the curve were doubled?

Solution

The curve would be farther to the right and higher up, but the same basic shape.

25. Determine the volume of 1 mol of CH₄ gas at 150 K and 1 atm, using Figure 9.12.

Solution

The figure shows the change in volume for 1 mol of CH₄ gas as a function of temperature. The graph shows that the volume is about 12.5 L.

27. A spray can is used until it is empty except for the propellant gas, which has a pressure of 1344 torr at 23 °C. If the can is thrown into a fire ($T = 475$ °C), what will be the pressure in the hot can?

Solution

The first thing to recognize about this problem is that the volume and moles of gas remain constant. Thus, we can use the combined gas law equation in the form:

$$\frac{P_2}{T_2} = \frac{P_1}{T_1}$$

$$P_2 = \frac{P_1 T_2}{T_1} = 1344 \text{ torr} \times \frac{475 + 273.15}{23 + 273.15} = 3.40 \times 10^3 \text{ torr}$$

29. A 2.50-L volume of hydrogen measured at -196 °C is warmed to 100 °C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

Solution

Apply Charles's law to compute the volume of gas at the higher temperature:

$$V_1 = 2.50 \text{ L}$$

$$T_1 = -196 \text{ °C} = 77.15 \text{ K}$$

$$V_2 = ?$$

$$T_2 = 100 \text{ °C} = 373.15 \text{ K}$$

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

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{2.50 \text{ L} \times 373.15 \text{ K}}{77.15 \text{ K}} = 12.1 \text{ L}$$

31. A weather balloon contains 8.80 moles of helium at a pressure of 0.992 atm and a temperature of 25° C at ground level. What is the volume of the balloon under these conditions?



Solution

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{8.80 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298.15 \text{ K}}{0.992 \text{ atm}} = 217 \text{ L}$$

33. How many moles of gaseous boron trifluoride, BF_3 , are contained in a 4.3410-L bulb at 788.0 K if the pressure is 1.220 atm? How many grams of BF_3 ?

Solution

$$n = \frac{PV}{RT} = \frac{1.220 \text{ atm} (4.3410 \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(788.0 \text{ K})} = 0.08190 \text{ mol} = 8.190 \times 10^{-2} \text{ mol}$$

$$n \times \text{molar mass} = 8.190 \times 10^{-2} \text{ mol} \times 67.8052 \text{ g mol}^{-1} = 5.553 \text{ g}$$

35. How many grams of gas are present in each of the following cases?

- (a) 0.100 L of CO_2 at 307 torr and 26 °C
 (b) 8.75 L of C_2H_4 , at 378.3 kPa and 483 K
 (c) 221 mL of Ar at 0.23 torr and -54 °C

Solution

In each of these problems, we are given a volume, pressure, and temperature. We can obtain moles from this information using the molar mass, $m = nM$, where M is the molar mass:

$$P, V, T \xrightarrow{n = PV/RT} n, \xrightarrow{m = n (\text{molar mass})} \text{grams}$$

or we can combine these equations to obtain:

$$\text{mass} = m = \frac{PV}{RT} \times M$$

(a)

$$307 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.4039 \text{ atm} \quad 26 \text{ }^\circ\text{C} = 299.1 \text{ K}$$

$$\text{Mass} = m = \frac{0.4039 \text{ atm} (0.100 \text{ L})}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} (299.1 \text{ K})} \times 44.01 \text{ g mol}^{-1} = 7.24 \times 10^{-2} \text{ g}$$

(b)

$$\text{Mass} = m = \frac{378.3 \text{ kPa} (8.75 \text{ L})}{8.314 \text{ L kPa mol}^{-1} \text{ K}^{-1} (483 \text{ K})} \times 28.05376 \text{ g mol}^{-1} = 23.1 \text{ g}$$

(c)

$$221 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.221 \text{ L} \quad -54 \text{ }^\circ\text{C} + 273.15 = 219.15 \text{ K}$$

$$0.23 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 3.03 \times 10^{-4} \text{ atm}$$

$$\text{Mass} = m = \frac{3.03 \times 10^{-4} \text{ atm} (0.221 \text{ L})}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} (219.15 \text{ K})} \times 39.948 \text{ g mol}^{-1} = 1.5 \times 10^{-4} \text{ g}$$

37. A cylinder of medical oxygen has a volume of 35.4 L, and contains O₂ at a pressure of 151 atm and a temperature of 25 °C. What volume of O₂ does this correspond to at normal body conditions, that is, 1 atm and 37 °C?

Solution

$$\begin{aligned} \frac{P_1 V_1}{T_1} &= \frac{P_2 V_2}{T_2} \\ V_2 &= \frac{P_1 V_1 T_2}{T_1 P_2} \\ &= \frac{(151 \text{ atm})(35.4 \text{ L})(310 \text{ K})}{(298 \text{ K})(1 \text{ atm})} = 5561 \text{ L} \end{aligned}$$

39. A 20.0-L cylinder containing 11.34 kg of butane, C₄H₁₀, was opened to the atmosphere. Calculate the mass of the gas remaining in the cylinder if it were opened and the gas escaped until the pressure in the cylinder was equal to the atmospheric pressure, 0.983 atm, and a temperature of 27 °C.

Solution

Calculate the amount of butane in 20.0 L at 0.983 atm and 27 °C. The original amount in the container does not matter.

$$n = \frac{PV}{RT} = \frac{0.983 \text{ atm} \times 20.0 \text{ L}}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} (300.1 \text{ K})} = 0.798 \text{ mol}$$

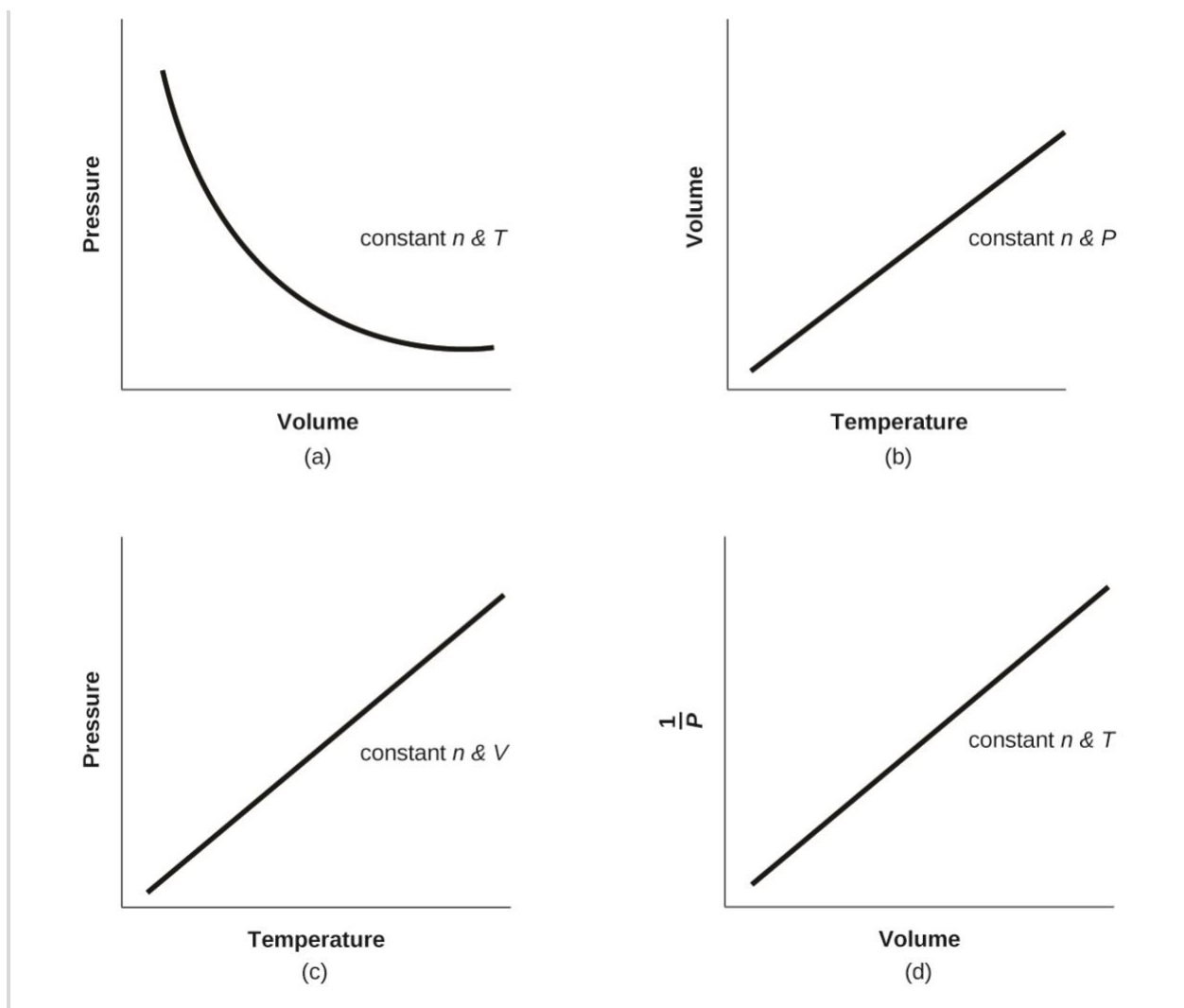
$$\text{Mass of butane} = 0.798 \text{ mol} \times 58.1234 \text{ g/mol} = 46.4 \text{ g}$$

41. For a given amount of gas showing ideal behavior, draw labeled graphs of:

- the variation of P with V
- the variation of V with T
- the variation of P with T
- the variation of $\frac{1}{P}$ with V

Solution

For a gas exhibiting ideal behavior:



43. The effect of chlorofluorocarbons (such as CCl_2F_2) on the depletion of the ozone layer is well known. The use of substitutes, such as $\text{CH}_3\text{CH}_2\text{F}(g)$, for the chlorofluorocarbons, has largely corrected the problem. Calculate the volume occupied by 10.0 g of each of these compounds at STP:

(a) $\text{CCl}_2\text{F}_2(g)$

(b) $\text{CH}_3\text{CH}_2\text{F}(g)$

Solution

(a) Determine the molar mass of CCl_2F_2 then calculate the moles of $\text{CCl}_2\text{F}_2(g)$ present. Use the ideal gas law $PV = nRT$ to calculate the volume of $\text{CCl}_2\text{F}_2(g)$:

$$10.0 \text{ g CCl}_2\text{F}_2 \times \frac{1 \text{ mol CCl}_2\text{F}_2}{120.91 \text{ g CCl}_2\text{F}_2} = 0.0827 \text{ mol CCl}_2\text{F}_2$$

$PV = nRT$, where $n = \# \text{ mol CCl}_2\text{F}_2$

$$1 \text{ atm} \times V = 0.0827 \text{ mol} \times \frac{0.0821 \text{ L atm}}{\text{mol K}} \times 273 \text{ K} = 1.85 \text{ L CCl}_2\text{F}_2;$$

$$(b) 10.0 \text{ g CH}_3\text{CH}_2\text{F} \times \frac{1 \text{ mol CH}_3\text{CH}_2\text{F}}{48.07 \text{ g CH}_3\text{CH}_2\text{F}} = 0.208 \text{ mol CH}_3\text{CH}_2\text{F}$$

$PV = nRT$, with $n = \# \text{ mol CH}_3\text{CH}_2\text{F}$

$$1 \text{ atm} \times V = 0.208 \text{ mol} \times 0.0821 \text{ L atm/mol K} \times 273 \text{ K} = 4.66 \text{ L CH}_3\text{CH}_2\text{F}$$

45. A balloon with a volume of 100.21 L at 21 °C and 0.981 atm is released and just barely clears the top of Mount Crumpet in British Columbia. If the final volume of the balloon is 144.53 L at a temperature of 5.24 °C, what is the pressure experienced by the balloon as it clears Mount Crumpet?

Solution

Identify the variables in the problem and determine that the combined gas law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ is the necessary equation to use to solve the problem. Then solve for P_2 :

$$\frac{0.981 \text{ atm} \times 100.21 \text{ L}}{294 \text{ K}} = \frac{P_2 \times 144.53 \text{ L}}{278.24 \text{ K}}$$

$$P_2 = 0.644 \text{ atm}$$

47. If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

Solution

The pressure decreases by a factor of 3.

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Chemistry 2e

9: Gases

9.3: Stoichiometry of Gaseous Substances, Mixtures, and Reactions

49. Calculate the density of Freon 12, CF_2Cl_2 , at 30.0°C and 0.954 atm .

Solution

$$\rho = \frac{PM}{RT} = \frac{0.954\text{ atm} [12.011 + 2(18.9954) + 2(35.453)]\text{ g mol}^{-1}}{0.08206\text{ L atm mol}^{-1}\text{ K}^{-1} \times 303.15\text{ K}} = 4.64\text{ g L}^{-1}$$

51. A cylinder of $\text{O}_2(\text{g})$ used in breathing by patients with emphysema has a volume of 3.00 L at a pressure of 10.0 atm . If the temperature of the cylinder is 28.0°C , what mass of oxygen is in the cylinder?

Solution

$$\text{mass O}_2 = \frac{(31.9988\text{ g mol}^{-1})(10.0\text{ atm})(3.00\text{ L})}{(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(301.15\text{ K})} = 38.8\text{ g}$$

53. What is the molar mass of a gas if 0.281 g of the gas occupies a volume of 125 mL at a temperature 126°C and a pressure of 777 torr ?

Solution

From the ideal gas law, $PV = nRT$, set $n = \frac{\text{mass}}{\text{molar mass}}$ and solve the molar mass.

$$\text{molar mass} = \frac{(0.281\text{ g})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(399.15\text{ K})}{\left(\frac{777\text{ torr}}{760\text{ torr atm}^{-1}}\right)(0.125\text{ L})} = 72.0\text{ g mol}^{-1}$$

55. The density of a certain gaseous fluoride of phosphorus is 3.93 g/L at STP. Calculate the molar mass of this fluoride and determine its molecular formula.

Solution

$$M = \frac{mRT}{PV} \quad D = \frac{m}{V} \quad M = \frac{DRT}{P}$$

$$M = \frac{3.93\text{ g L}^{-1} (0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(273.15\text{ K})}{1.00\text{ atm}} = 88.1\text{ g mol}^{-1}$$

$$M_{\text{phosphorous}} = 30.97376\text{ g/mol}$$

$$M_{\text{fluorine}} = 18.998403\text{ g/mol}$$

molecular formula: phosphorous: 30.97376

$$\text{fluorine: } \frac{3(18.998403)}{87.968969}$$

The molecular formula is PF_3 .

To find this answer you can either use trial and error, or you can realize that since phosphorus is in group 5, it can fill its valence shell by forming three bonds. Fluorine, being in group 7, needs to form only one bond to fill its shell. Thus it makes sense to start with PF_3 as a probable formula.

57. A 36.0-L cylinder of a gas used for calibration of blood gas analyzers in medical laboratories contains 350 g CO_2 , 805 g O_2 , and $4,880\text{ g N}_2$. At 25°C , what is the pressure in the cylinder in atmospheres, in torr, and in kilopascals?

Solution

Calculate the moles of each gas present and from that, calculate the pressure from the ideal gas law. Assume 25 °C. The calibration gas contains:

$$\frac{350 \text{ g CO}_2}{44.0098 \text{ g mol}^{-1} \text{ CO}_2} = 7.953 \text{ mol CO}_2$$

$$\frac{805 \text{ g O}_2}{31.9988 \text{ g mol}^{-1} \text{ O}_2} = 25.157 \text{ mol O}_2$$

$$\frac{4880 \text{ g N}_2}{28.01348 \text{ g mol}^{-1} \text{ N}_2} = 174.202 \text{ mol N}_2$$

Total moles = 7.953 + 25.157 + 174.202 = 207.312 mol

$$P = \frac{nRT}{V} = \frac{207.312 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298.15 \text{ K}}{36.0 \text{ L}} = 141 \text{ atm}$$

P in torr = 107,000 torr

P in kPa = 14,300 kPa

59. A sample of gas isolated from unrefined petroleum contains 90.0% CH₄, 8.9% C₂H₆, and 1.1% C₃H₈ at a total pressure of 307.2 kPa. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

Solution

Since these are percentages of the total pressure, the partial pressure can be calculated as follows:

CH₄: 90% of 307.2 kPa = 0.900 × 307.2 = 276 kPa

C₂H₆: 8.9% of 307.2 kPa = 0.089 × 307.2 = 27 kPa

C₃H₈: 1.1% of 307.2 kPa = 0.011 × 307.2 = 3.4 kPa

61. Most mixtures of hydrogen gas with oxygen gas are explosive. However, a mixture that contains less than 3.0 % O₂ is not. If enough O₂ is added to a cylinder of H₂ at 33.2 atm to bring the total pressure to 34.5 atm, is the mixture explosive?

Solution

The oxygen increases the pressure within the tank to (34.5 atm – 33.2 atm =) 1.3 atm. The

percentage O₂ on a mole basis is $\frac{1.3}{34.5} \times 100\% = 3.77\%$. The mixture is explosive. However, the

percentage is given as a weight percent. Converting to a mass basis increases the percentage of oxygen even more, so the mixture is still explosive.

63. A sample of carbon monoxide was collected over water at a total pressure of 756 torr and a temperature of 18 °C. What is the pressure of the carbon monoxide? (See Table 9.2 for the vapor pressure of water.)

Solution

The vapor pressure of water at 18 °C is 15.5 torr. Subtract the vapor pressure of water from the total pressure to find the pressure of the carbon monoxide:

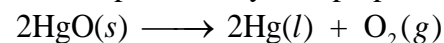
$$P_T = P_{\text{gas}} + P_{\text{water}}$$

Rearrangement gives:

$$P_T - P_{\text{water}} = P_{\text{gas}}$$

$$756 \text{ torr} - 15.5 \text{ torr} = 740 \text{ torr}$$

65. Joseph Priestley first prepared pure oxygen by heating mercuric oxide, HgO:



(a) Outline the steps necessary to answer the following question: What volume of O₂ at 23 °C and 0.975 atm is produced by the decomposition of 5.36 g of HgO?

(b) Answer the question.

Solution

(a) Determine the moles of HgO that decompose; using the chemical equation, determine the moles of O₂ produced by decomposition of this amount of HgO; and determine the volume of O₂ from the moles of O₂, temperature, and pressure.

(b)

$$5.36 \text{ g HgO} \times \frac{1 \text{ mol HgO}}{(200.59 + 15.9994) \text{ g HgO}} = 0.0247 \text{ mol HgO}$$

$$0.0247 \text{ mol HgO} \times \frac{1 \text{ mol O}_2}{2 \text{ mol HgO}} = 0.01235 \text{ mol O}_2$$

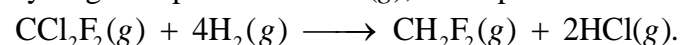
$$PV = nRT$$

$$P = 0.975 \text{ atm}$$

$$T = (23.0 + 273.15) \text{ K}$$

$$V = \frac{nRT}{P} = \frac{0.01235 \text{ mol} (0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(296.15 \text{ K})}{0.975 \text{ atm}} = 0.308 \text{ L}$$

67. The chlorofluorocarbon CCl₂F₂ can be recycled into a different compound by reaction with hydrogen to produce CH₂F₂(g), a compound useful in chemical manufacturing:



(a) Outline the steps necessary to answer the following question: What volume of hydrogen at 225 atm and 35.5 °C would be required to react with 1 ton (1.000 × 10³ kg) of CCl₂F₂?

(b) Answer the question.

Solution

(a) Determine the molar mass of CCl₂F₂. From the balanced equation, calculate the moles of H₂ needed for the complete reaction. From the ideal gas law, convert moles of H₂ into volume.

(b) Molar mass of CCl₂F₂ = 12.011 + 2 × 18.9984 + 2 × 35.4527 = 120.913 g/mol

$$\text{mol H}_2 = 1.000 \times 10^6 \text{ g} \times \frac{1 \text{ mol CCl}_2\text{F}_2}{120.913 \text{ g}} \times \frac{4 \text{ mol H}_2}{1 \text{ mol CCl}_2\text{F}_2} = 3.308 \times 10^4 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(3.308 \times 10^4 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(308.65 \text{ K})}{225 \text{ atm}} = 3.72 \times 10^3 \text{ L}$$

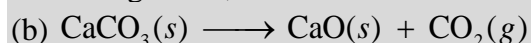
69. Lime, CaO, is produced by heating calcium carbonate, CaCO₃; carbon dioxide is the other product.

(a) Outline the steps necessary to answer the following question: What volume of carbon dioxide at 875 K and 0.966 atm is produced by the decomposition of 1 ton (1.000 × 10³ kg) of calcium carbonate?

(b) Answer the question.

Solution

(a) Balance the equation. Determine the grams of CO₂ produced and the number of moles. From the ideal gas law, determine the volume of gas.



$$\text{mass CO}_2 = 1.00 \times 10^6 \text{ g} \times \frac{1 \text{ mol CaCO}_3}{100.087 \text{ g}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} = 4.397 \times 10^5 \text{ g}$$

$$\text{mol CO}_2 = \frac{4.397 \times 10^5 \text{ g}}{44.01 \text{ g mol}^{-1}} = 9991 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(9991 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(875 \text{ K})}{0.966 \text{ atm}} = 7.43 \times 10^5 \text{ L}$$

71. Calculate the volume of oxygen required to burn 12.00 L of ethane gas, C_2H_6 , to produce carbon dioxide and water, if the volumes of C_2H_6 and O_2 are measured under the same conditions of temperature and pressure.

Solution



From the balanced equation, we see that 2 mol of C_2H_6 requires 7 mol of O_2 to burn completely. Gay-Lussac's law states that gases react in simple proportions by volume. As the number of liters is proportional to the number of moles,

$$\frac{12.00 \text{ L}}{2 \text{ mol C}_2\text{H}_6} = \frac{V(\text{O}_2)}{7 \text{ mol O}_2}$$

$$V(\text{O}_2) = \frac{12.00 \text{ L} \times 7}{2} = 42.00 \text{ L}$$

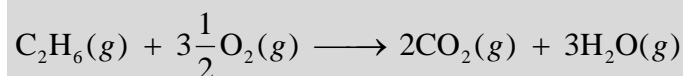
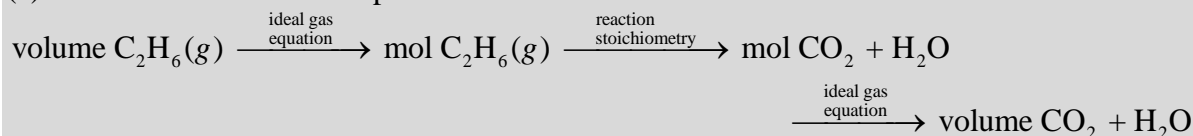
73. Consider the following questions:

(a) What is the total volume of the $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ at 600°C and 0.888 atm produced by the combustion of 1.00 L of $\text{C}_2\text{H}_6(\text{g})$ measured at STP?

(b) What is the partial pressure of H_2O in the product gases?

Solution

(a) The scheme to solve this problem is:



$$1. n(\text{C}_2\text{H}_6) = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 1.00 \text{ L}}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}(273.15 \text{ K})} = 0.0446 \text{ mol}$$

$$2. 0.0446 \text{ mol C}_2\text{H}_6 \times \frac{5 \text{ mol products}}{1 \text{ mol C}_2\text{H}_6} = 0.223 \text{ mol products}$$

$$3. V = nRT = \frac{(0.223 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(873.15 \text{ K})}{0.888 \text{ atm}} = 18.0 \text{ L}$$

(b) First, calculate the mol H_2O produced:

$$0.0446 \text{ mol C}_2\text{H}_6 \times \frac{3 \text{ mol products}}{1 \text{ mol C}_2\text{H}_6} = 0.1338 \text{ mol}$$

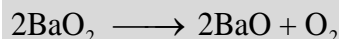
Second, calculate the pressure of H_2O :

$$P = \frac{nRT}{V} = \frac{(0.1338 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(873.15 \text{ K})}{18.0 \text{ L}} = 0.533 \text{ atm}$$

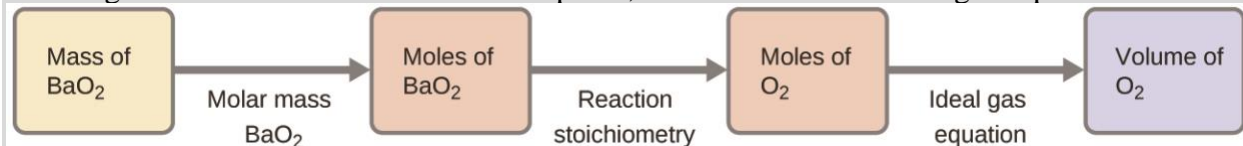
75. What volume of oxygen at 423.0 K and a pressure of 127.4 kPa is produced by the decomposition of 129.7 g of BaO_2 to BaO and O_2 ?

Solution

First, we must write a balanced equation to establish the stoichiometry of the reaction:



We are given the mass of BaO_2 that decomposes, so the scheme for solving this problem will be:

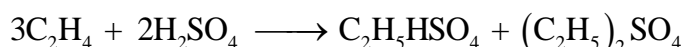


$$\text{Mass}(\text{BaO}_2) = 137.33 + 2(15.9994) = 169.33 \text{ g/mol}$$

$$n(\text{O}_2) = 129.7 \text{ g BaO}_2 \times \frac{1 \text{ mol BaO}_2}{169.33 \text{ g BaO}_2} \times \frac{1 \text{ mol O}_2}{2 \text{ mol BaO}_2} = 0.3830 \text{ mol O}_2$$

$$V(\text{O}_2) = \frac{nRT}{P} = \frac{0.3830 \text{ mol}(8.314 \text{ L kPa mol}^{-1} \text{ K}^{-1})(423.0 \text{ K})}{127.4 \text{ kPa}} = 10.57 \text{ L O}_2$$

77. Ethanol, $\text{C}_2\text{H}_5\text{OH}$, is produced industrially from ethylene, C_2H_4 , by the following sequence of reactions:



What volume of ethylene at STP is required to produce 1.000 metric ton (1000 kg) of ethanol if the overall yield of ethanol is 90.1%?

Solution

At 90.1% conversion, a $1.000 \times 10^6 \text{ g}$ final yield would require a $\left(\frac{1.000 \times 10^6}{0.901}\right) = 1.1099 \times 10^6 \text{ g}$

theoretical yield.

$3\text{C}_2\text{H}_4$ produces $3\text{C}_2\text{H}_5\text{OH}$, giving a 1:1 ratio:

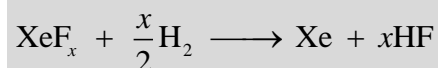
$$\begin{aligned} \text{mol}(\text{C}_2\text{H}_4) &= 1.1099 \times 10^6 \text{ g } \cancel{\text{C}_2\text{H}_5\text{OH}} \times \frac{1 \text{ mol } \cancel{\text{C}_2\text{H}_2\text{OH}}}{46.069 \text{ g } \cancel{\text{C}_2\text{H}_5\text{OH}}} \times \frac{1 \text{ mol } \text{C}_2\text{H}_4}{1 \text{ mol } \cancel{\text{C}_2\text{H}_2\text{OH}}} \\ &= 2.409 \times 10^4 \text{ mol} \end{aligned}$$

$$V(\text{C}_2\text{H}_4) = 22.4 \text{ L/mol} \times 2.409 \times 10^4 \text{ mol} = 5.40 \times 10^5 \text{ L}$$

79. A sample of a compound of xenon and fluorine was confined in a bulb with a pressure of 18 torr. Hydrogen was added to the bulb until the pressure was 72 torr. Passage of an electric spark through the mixture produced Xe and HF. After the HF was removed by reaction with solid KOH, the final pressure of xenon and unreacted hydrogen in the bulb was 36 torr. What is the empirical formula of the xenon fluoride in the original sample? (Note: Xenon fluorides contain only one xenon atom per molecule.)

Solution

The reaction is:



Immediately after the H_2 is added (before the reaction):

$$P_{\text{Total}} = P_{\text{XeF}_2} + P_{\text{H}_2}$$

$$P_{\text{H}_2} = P_{\text{Total}} - P_{\text{XeF}_2}$$

$$= 72 \text{ torr} - 18 \text{ torr}$$

$$= 54 \text{ torr}$$

After the reaction:

$$P_{\text{Xe}} = 18 \text{ torr} \quad (1 \text{ mol XeF}_x \longrightarrow 1 \text{ mol Xe})$$

And the partial pressure of unreacted H₂ is:

$$P_{\text{H}_2} = P_{\text{Total}} - P_{\text{Xe}}$$

$$= 36 \text{ torr} - 18 \text{ torr}$$

$$= 18 \text{ torr}$$

The pressure of H₂ that reacts is:

$$54 \text{ torr} - 18 \text{ torr} = 36 \text{ torr}$$

The number of moles of gas is proportional to the partial pressures. The reaction used 18 torr of XeF_x and 36 torr of H₂ so:

$$\frac{\text{mol H}_2}{\text{mol XeF}_x} = \frac{x/2}{1} = \frac{x}{2} = \frac{36 \text{ torr}}{18 \text{ torr}} \longrightarrow x = \frac{72 \text{ torr}}{18 \text{ torr}} = 4$$

The empirical formula for the compound is XeF₄.

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Chemistry 2e
9: Gases
9.4: Effusion and Diffusion of Gases

81. A balloon filled with helium gas takes 6 hours to deflate to 50% of its original volume. How long will it take for an identical balloon filled with the same volume of hydrogen gas (instead of helium) to decrease its volume by 50%?

Solution

Use the rate of effusion equation:

$$\frac{6 \text{ hours}}{t} = \sqrt{\frac{4}{2}}$$

$$t = \frac{6 \text{ hours}}{1.4} = 4.2 \text{ hours}$$

83. Starting with the definition of rate of effusion and Graham's finding relating rate and molar mass, show how to derive the Graham's law equation, relating the relative rates of effusion for two gases to their molecular masses.

Solution

Effusion can be defined as the process by which a gas escapes through a pinhole into a vacuum. Graham's law states that with a mixture of two gases A and B:

$$\left(\frac{\text{rate A}}{\text{rate B}}\right) = \left(\frac{\text{molar mass of B}}{\text{molar mass of A}}\right)^{1/2}. \text{ Both A and B are in the same container at the same}$$

temperature and, therefore, will have the same kinetic energy:

$$KE_A = KE_B$$

$$KE = \frac{1}{2}mv^2$$

$$\text{Therefore, } \frac{1}{2}m_A v_A^2 = \frac{1}{2}m_B v_B^2$$

$$\frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\left(\frac{v_A^2}{v_B^2}\right)^{1/2} = \left(\frac{m_B}{m_A}\right)^{1/2}$$

$$\frac{v_A}{v_B} = \left(\frac{m_B}{m_A}\right)^{1/2}$$

85. Which of the following gases diffuse more slowly than oxygen? F₂, Ne, N₂O, C₂H₂, NO, Cl₂, H₂S

Solution

Gases with molecular masses greater than that of oxygen (31.9988 g/mol) will diffuse more slowly than O₂. These gases are F₂ (37.9968 g/mol), N₂O (44.0128 g/mol), Cl₂ (70.906 g/mol), and H₂S (34.082 g/mol).

87. Calculate the relative rate of diffusion of ¹H₂ (molar mass 2.0 g/mol) compared with that of ²H₂ (molar mass 4.0 g/mol) and the relative rate of diffusion of O₂ (molar mass 32 g/mol) compared to that of O₃ (molar mass 48 g/mol).

Solution

$$\frac{R_{\text{H}_2}}{R_{\text{D}_2}} = \frac{\sqrt{M_{\text{D}_2}}}{\sqrt{M_{\text{H}_2}}} = \frac{\sqrt{4.0}}{\sqrt{2.0}} = \frac{2.000}{1.414} = 1.4$$

$$\frac{R_{\text{O}_2}}{R_{\text{O}_3}} = \frac{\sqrt{M_{\text{O}_3}}}{\sqrt{M_{\text{O}_2}}} = \frac{\sqrt{48}}{\sqrt{32}} = 1.2$$

89. When two cotton plugs, one moistened with ammonia and the other with hydrochloric acid, are simultaneously inserted into opposite ends of a glass tube that is 87.0 cm long, a white ring of NH_4Cl forms where gaseous NH_3 and gaseous HCl first come into contact. (Hint: Calculate the rates of diffusion for both NH_3 and HCl , and find out how much faster NH_3 diffuses than HCl .)

$\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \longrightarrow \text{NH}_4\text{Cl}(\text{s})$ At approximately what distance from the ammonia-moistened plug does this occur?

Solution

Rate of diffusion for NH_3 is proportional to $\frac{1}{17.04^{1/2}} = 0.242250792$

Rate of diffusion for HCl is proportional to $\frac{1}{36.46^{1/2}} = 0.165611949$, $\left(\frac{0.242250792}{0.165611949}\right) = 1.4627$.

Set up an algebraic expression, letting x represent the distance travelled by the HCl : $x + 1.4627x = 87$, $x = 35.3$, so the distance travelled by the NH_3 is $(1.4627)x = 51.7$ cm.

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Chemistry 2e
9: Gases
9.5: The Kinetic-Molecular Theory

91. Can the speed of a given molecule in a gas double at constant temperature? Explain your answer.

Solution

Yes. At any given instant, there are a range of values of molecular speeds in a sample of gas. Any single molecule can speed up or slow down as it collides with other molecules. The average speed of all the molecules is constant at constant temperature.

93. The distribution of molecular speeds in a sample of helium is shown in Figure 9.34. If the sample is cooled, will the distribution of speeds look more like that of H₂ or of H₂O? Explain your answer.

Solution

H₂O. Cooling slows the speeds of the He atoms, causing them to behave as though they were heavier.

95. A 1-L sample of CO initially at STP is heated to 546 K, and its volume is increased to 2 L.
- What effect do these changes have on the number of collisions of the molecules of the gas per unit area of the container wall?
 - What is the effect on the average kinetic energy of the molecules?
 - What is the effect on the root mean square speed of the molecules?

Solution

Both the temperature and the volume are doubled for this gas (n constant), so P remains constant.

(a) The number of collisions per unit area of the container wall is constant. (b) The average kinetic energy doubles; it is proportional to temperature. (c) The root mean square speed increases to $\sqrt{2}$ times its initial value; u_{rms} is proportional to $\sqrt{\text{KE}_{\text{avg}}}$.

97. Answer the following questions:

- Is the pressure of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?
- Is the density of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?
- At a pressure of 1 atm and a temperature of 20 °C, dry air has a density of 1.2256 g/L. What is the (average) molar mass of dry air?
- The average temperature of the gas in a hot-air balloon is 1.30×10^2 °F. Calculate its density, assuming the molar mass equals that of dry air.
- The lifting capacity of a hot-air balloon is equal to the difference in the mass of the cool air displaced by the balloon and the mass of the gas in the balloon. What is the difference in the mass of 1.00 L of the cool air in part (c) and the hot air in part (d)?
- An average balloon has a diameter of 60 feet and a volume of 1.1×10^5 ft³. What is the lifting power of such a balloon? If the weight of the balloon and its rigging is 500 pounds, what is its capacity for carrying passengers and cargo?
- A balloon carries 40.0 gallons of liquid propane (density 0.5005 g/L). What volume of CO₂ and H₂O gas is produced by the combustion of this propane?
- A balloon flight can last about 90 minutes. If all of the fuel is burned during this time, what is the approximate rate of heat loss (in kJ/min) from the hot air in the bag during the flight?

Solution

(a) equal, because the balloon is free to expand until the pressures are equalized; (b) less than the density outside; (c) assume three-place accuracy throughout unless greater accuracy is stated:

$$\text{molar mass} = \frac{DRT}{P} = 1.2256 \text{ g L}^{-1} \times \frac{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 293.15 \text{ K}}{1.00 \text{ atm}};$$

$$= 29.48 \text{ g mol}^{-1}$$

(d) convert the temperature to °C; then use the ideal gas law:

$$^{\circ}\text{C} = \frac{5}{9}(\text{F} - 32) = \frac{5}{9}(130 - 32) = 54.44 \text{ }^{\circ}\text{C} = 327.6 \text{ K}$$

$$D = \frac{MP}{RT} = 29.48 \text{ g mol}^{-1} \times \frac{1.00 \text{ atm}}{0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 327.6 \text{ K}} = 1.0966 \text{ g L}^{-1};$$

(e) $1.2256 \text{ g/L} - 1.09966 \text{ g/L} = 0.129 \text{ g/L}$; (f) calculate the volume in liters, multiply the volume by the density difference to find the lifting capacity of the balloon, subtract the weight of the balloon after converting to pounds:

$$1.1 \times 10^5 \text{ ft}^3 \times \left(\frac{12 \text{ in}}{91 \text{ ft}}\right)^3 \times \left(\frac{2.54 \text{ cm}}{\text{in}}\right)^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 3.11 \times 10^6 \text{ L}$$

$$3.11 \times 10^6 \text{ L} \times 0.129 \text{ g/L} = 4.01 \times 10^5 \text{ g}$$

$$\frac{4.01 \times 10^5 \text{ g}}{453.59 \text{ g lb}^{-1}} = 884 \text{ lb}; 884 \text{ lb} - 500 \text{ lb} = 384 \text{ lb}$$

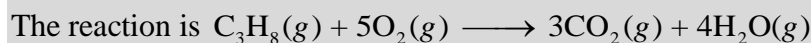
net lifting capacity = 384 lb; (g) First, find the mass of propane contained in 40.0 gal. Then calculate the moles of $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ produced from the balanced equation.

$$40.0 \text{ gal} \times \frac{4(0.9463 \text{ L})}{1 \text{ gal}} = 151.4 \text{ L}$$

$$151.4 \text{ L} \times 0.5005 \text{ g L}^{-1} = 75.8 \text{ g}$$

$$\text{Molar mass of propane} = 3(12.011) + 8(1.00794) = 36.033 + 8.064 = 44.097 \text{ g mol}^{-1}$$

$$\frac{75.8 \text{ g}}{44.097 \text{ g mol}^{-1}} = 1.72 \text{ mol}$$



For each 1.72 mol propane, there are $3 \times 1.72 \text{ mol} = 5.15 \text{ mol}$ of CO_2 and $4 \times 1.72 \text{ mol} = 6.88 \text{ mol}$ H_2O . The total volume at STP = $22.4 \text{ L} \times 12.04 = 270 \text{ L}$; (h) The total heat released is determined from the heat of combustion of the propane. Using the equation in part (g),

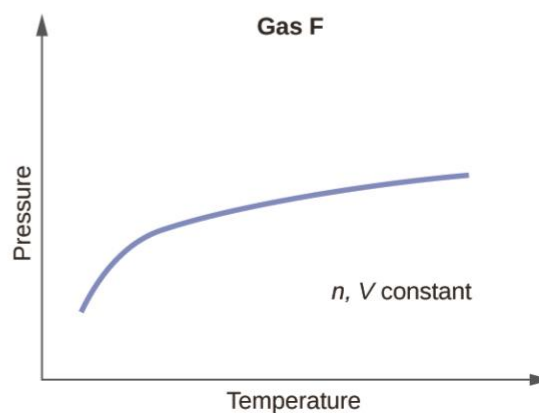
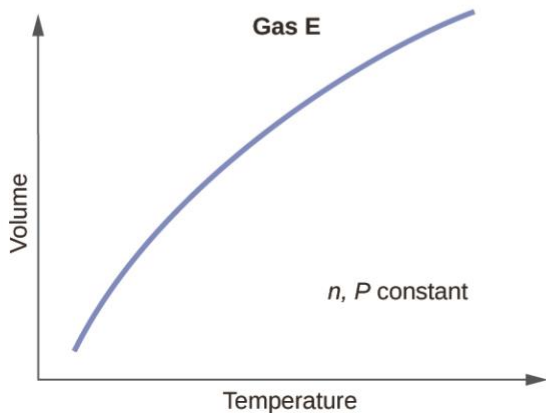
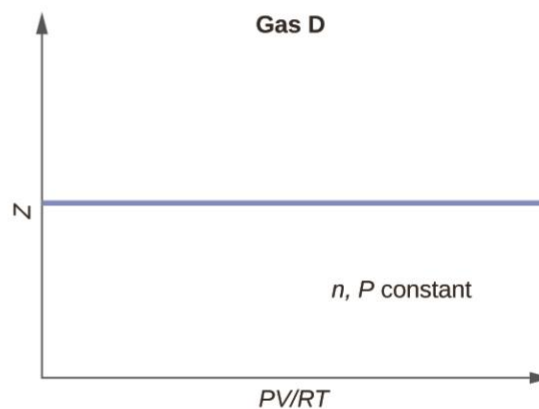
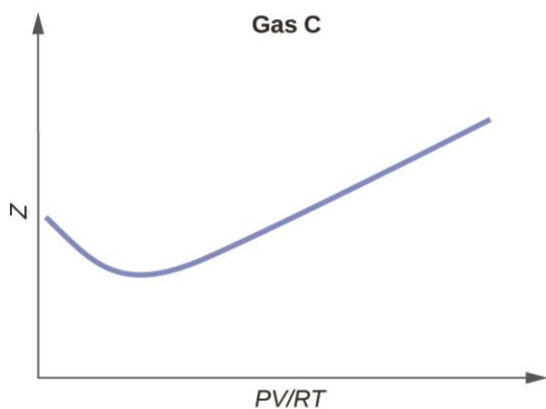
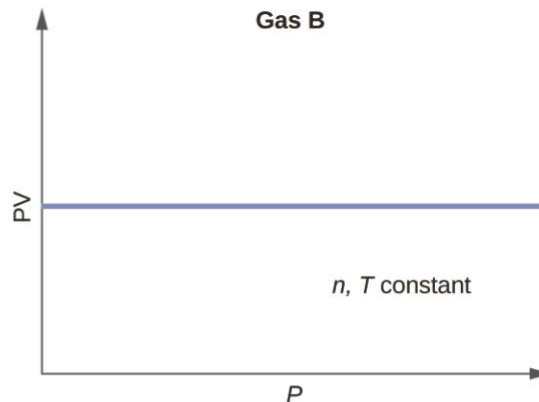
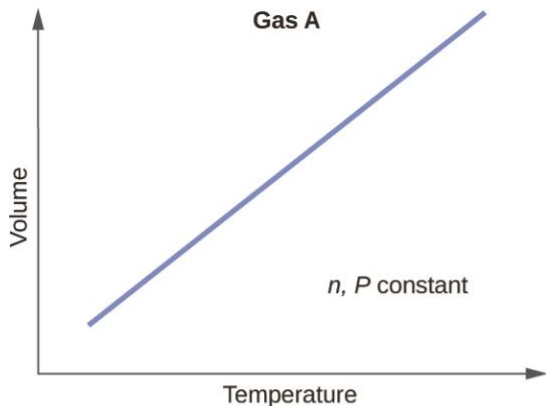
$$\begin{aligned} \Delta H_{\text{combustion}}^{\circ} &= 3\Delta H_{\text{CO}_2(\text{g})}^{\circ} + 4\Delta H_{\text{H}_2\text{O}(\text{g})}^{\circ} - \Delta H_{\text{propane}}^{\circ} \\ &= 3(-393.51) + 4(-241.82) - (-103.85) \\ &= -1180.52 - 967.28 + 103.85 = -2043.96 \text{ kJ mol}^{-1} \end{aligned}$$

Since there is 1.72 mol propane, $1.72 \times 2043.96 \text{ kJ mol}^{-1} = 3.52 \times 10^3 \text{ kJ}$ is used for heating.

This heat is used over 90 minutes, so $\frac{3.52 \times 10^3 \text{ kJ}}{90 \text{ min}} = 39.1 \text{ kJ min}^{-1}$ is released.

Chemistry 2e
9: Gases
9.6: Non-Ideal Gas Behavior

99. Graphs showing the behavior of several different gases follow. Which of these gases exhibit behavior significantly different from that expected for ideal gases?



Solution

Gas A: volume increases linearly as temperature increases with moles and pressure held constant, as expected by the ideal gas law $V = (nR/P)T$; Gas B: PV stays constant as pressure increases with moles and temperature held constant, as expected by the ideal gas law $PV = nRT$;

Gas C: compressibility factor (Z) varies as PV/RT increases, as expected of a real gas; Gas D: compressibility factor (Z) stays constant as PV/RT increases with moles and pressure held constant, as expected of an ideal gas; Gas E: as temperature increases, volume increases, but not linearly with moles and pressure held constant, as would *not* be expected by the ideal gas law $V = (nR/P)T$, as seen in Gas A; Gas F: as temperature increases, pressure increases with moles and volume held constant, but not linearly, as would *not* be expected by the ideal gas law $P = (nR/V)T$, as seen in Gas A; Gases C, E, and F exhibit behavior significantly different from that expected for an ideal gas.

101. Under which of the following sets of conditions does a real gas behave most like an ideal gas, and for which conditions is a real gas expected to deviate from ideal behavior? Explain.

- (a) high pressure, small volume
- (b) high temperature, low pressure
- (c) low temperature, high pressure

Solution

The gas behavior most like an ideal gas will occur under the conditions in (b). Molecules have high speeds and move through greater distances between collisions; they also have shorter contact times and interactions are less likely. Deviations occur with the conditions described in (a) and (c). Under conditions of (a), some gases may liquefy. Under conditions of (c), most gases will liquefy.

103. For which of the following gases should the correction for the molecular volume be largest: CO, CO₂, H₂, He, NH₃, SF₆?

Solution

We would expect the molecule with the largest volume to need the largest correction. SF₆ would need the largest correction.

105. Answer the following questions:

- (a) If XX behaved as an ideal gas, what would its graph of Z vs. P look like?
- (b) For most of this chapter, we performed calculations treating gases as ideal. Was this justified?
- (c) What is the effect of the volume of gas molecules on Z ? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.
- (d) What is the effect of intermolecular attractions on the value of Z ? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.
- (e) In general, under what temperature conditions would you expect Z to have the largest deviations from the Z for an ideal gas?

Solution

Answer: (a) A straight horizontal line at 1.0 (see Figure 9.35) for the line representing an ideal gas. (b) When real gases are at low pressures and high temperatures, they behave close enough to ideal gases that they are approximated as such; however, in some cases, we see that at a high pressure and temperature, the ideal gas approximation breaks down and is significantly different from the pressure calculated by the van der Waals equation. (c) The greater the compressibility, the more the volume matters. At low pressures, the correction factor for intermolecular attractions is more significant, and the effect of the volume of the gas molecules on Z would be a small lowering compressibility. At higher pressures, the effect of the volume of the gas molecules themselves on Z would increase compressibility (see Figure 9.35). (d) Once again, at low pressures, the effect of intermolecular attractions on Z would be more important than the correction factor for the volume of the gas molecules themselves, though perhaps still small. At

higher pressures and low temperatures, the effect of intermolecular attractions would be larger. See Figure 9.35. (e) Low temperatures

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