List the characteristics of an ideal gas.

An ideal gas has P-V-T behavior that is completely accounted for by the ideal gas equation, PV=nRT. Its molecules/atoms do not attract or repel each other (collisions are ideally elastic) and their volume is negligible compared to the volume of their container.

What are standard temperature and pressure? What is the significance of STP in relation to the volume of 1 mole of an ideal gas?

STP is 0 °C (273.15 K) and 1 atm (101.3 kPa). At STP, the molar volume of an ideal gas is 22.41 L.

Given 6.9 moles CO in 30.4 L at a temperature of 62°C, what is the pressure (in atm)?

\[ P = \frac{nRT}{V}, \text{ so } P = \frac{(6.9 \text{ mol})(0.08206 \text{ L atm/mol K})(62 + 273)K}{30.4 \text{ L}} = 6.2 \text{ atm} \]

A certain amount of gas at 25°C and at a pressure of 0.800 atm is contained in a glass vessel. Suppose that the vessel can withstand a pressure of 2.00 atm. How high can you raise the temperature of the gas without bursting the vessel?

\[ T_1 = (25 + 273)K = 298 \text{ K}; \ P_1 = 0.800 \text{ atm}; \ P_2 = 2.00 \text{ atm}; \ T_2 = ? \]

\[ \frac{P_1}{T_1} = \frac{P_2}{T_2} \Rightarrow T_2 = \frac{T_1 P_2}{P_1}, \text{ or } T_2 = \frac{(298 \text{ K})(2.00 \text{ atm})}{(0.800 \text{ atm})} = 745 \text{ K} = 472°C \]

A gas evolved during the fermentation of glucose has a volume of 0.78 L and 20.1 °C and 1.00 atm. What was the volume of this gas at the fermentation temperature of 36.5°C and 1.00 atm pressure?

\[ T_1 = (20.1 + 273) K = 293.1 \text{ K}; \ T_2 = (36.5 + 273)K = 309.5 \text{ K}; \ V_1 = 0.78 \text{ L}; \ V_2 = ? \]

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V_2 = \frac{V_1 T_2}{T_1}, \text{ or } V_2 = \frac{(0.78 \text{ L})(309.5 \text{ K})}{(293.1 \text{ K})} = 0.82 \text{ L} \]

Calculate the volume (in liters) of 88.4 g of CO₂ at STP.

\[ \mathcal{M} = 12.01 \text{ g/mol} + 2(16.00 \text{ g/mol}) = 44.01 \text{ g/mol} \]

\[ V_{\text{CO}_2} = 88.4 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 45.0 \text{ L CO}_2 \]
42. Dry ice is solid carbon dioxide. A 0.050g sample of dry ice is placed in an evacuated 4.6L vessel at 30°C. Calculate the pressure inside the vessel after all the dry ice has been converted to CO₂ gas.

\[ P = \frac{nRT}{V} \]

\[ P = \left( \frac{0.050 \text{ g} \times \frac{1 \text{ mol}}{44.01 \text{ g}}}{4.6 \text{ L}} \right) \left( 0.08206 \text{ L atm/mol K} \right) (30 + 273) \text{ K} = 6.1 \times 10^{-3} \text{ atm} \]

44. At 741 torr and 44°C, 7.10 g of a gas occupy a volume of 5.40 L. What is the molar mass of the gas?

\[ d = \frac{7.10 \text{ g}}{5.40 \text{ L}} = 1.31 \text{ g/L}, \quad T = 44° + 273° = 317 \text{ K}, \quad P = 741 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.975 \text{ atm} \]

Using \( PM = dRT \rightarrow \mathcal{M} = \frac{dRT}{P} \):

\[ \mathcal{M} = \frac{1.31 \text{ g/L}}{0.975 \text{ atm}} \left( 0.08206 \frac{\text{ L atm}}{\text{ mol K}} \right) (317 \text{ K}) = 35.0 \text{ g/mol} \]

48. Calculate the density of hydrogen bromide (HBr) gas in grams per liter at 733 mm Hg and 46°C.

\[ \mathcal{M} = 1.008 \text{ g/mol} + 79.90 \text{ g/mol} = 80.91 \text{ g/mol,} \quad T = 46° + 273° = 319 \text{ K} \]

Alternatively, we can solve for the density by writing: 

\[ \text{density} = \frac{\text{mass}}{\text{volume}} \]

If we have 1 mole of HBr, the mass is 80.91 g. Determine the volume:

\[ V = \frac{nRT}{P} \rightarrow V = \frac{(1 \text{ mol}) \left( 62.36 \frac{\text{L mmHg}}{\text{mol K}} \right) (319 \text{ K})}{733 \text{ mmHg}} = 27.2 \text{ L}, \quad \text{so} \quad D = \frac{\text{mass}}{\text{volume}} = \frac{80.91 \text{ g}}{27.2 \text{ L}} = 2.97 \text{ g/L} \]

50. A compound has the empirical formula of SF₄. At 20°C, 0.100 g of this gaseous compound occupies 22.1 mL at a pressure of 1.02 atm. What is the molecular formula?

We need to find molar mass: \( \mathcal{M} = \frac{dRT}{P} \);

\[ \mathcal{M} = \frac{0.100 \text{ g}}{0.0221 \text{ L}} \left( 0.08206 \frac{\text{ L atm}}{\text{ mol K}} \right) (20 + 273) \text{ K} = 107 \text{ g/mol} \]

emp. mass = 32.07 + 4(19.00) = 108.07 g/mol, so \( \frac{\text{molar mass}}{\text{empirical mass}} = 1 \) and mol. formula is SF₄.

54. In alcohol fermentation, yeast converts glucose to ethanol and carbon dioxide:

\[ \text{C}_6\text{H}_{12}\text{O}_6(s) \rightarrow 2 \text{C}_2\text{H}_5\text{OH}(l) + 2 \text{CO}_2(g) \]

If 5.97 g of glucose produce 1.44 L of CO₂ at 293 K and 0.984 atm, what is the % yield?

\[ \text{mol CO}_2 = 5.97 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.2 \text{ g glucose}} \times \frac{2 \text{ mol CO}_2}{1 \text{ mol glucose}} = 0.06626 \text{ mol CO}_2 \]

\[ V_{\text{CO}_2} = \frac{nRT}{P} = \frac{(0.06626 \text{ mol})(0.08206 \frac{\text{ L atm}}{\text{ mol K}})(293 \text{ K})}{0.984 \text{ atm}} = 1.619 \text{ L}; \quad \% \text{Yld} = \frac{1.44 \text{ L}}{1.619 \text{ L}} \times 100\% = 88.9\% \]
58 Reacting 3.00 g of an impure sample of calcium carbonate with hydrochloric acid produced 0.656 L of CO$_2$ (carbonic acid dissociates into carbon dioxide and water) at 20.0°C and 792 mmHg. Write the chemical equation and calculate the % by mass of calcium carbonate in the sample.

$$\text{CaCO}_3 (s) + 2 \text{HCl (aq)} \rightarrow \text{CaCl}_2 (aq) + \text{H}_2\text{O (l)} + \text{CO}_2 (g)$$

$$\text{mol CO}_2 = \frac{PV}{RT} = \frac{(792 \text{ mmHg})(0.656 \text{ L})}{(62.36 \frac{\text{L mmHg}}{\text{mol K}})(293 \text{ K})} = 0.02844 \text{ mol CO}_2$$

$$\text{mass CaCO}_3 = 0.02844 \text{ mol CO}_2 \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2} \times \frac{100.1 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 2.847 \text{ g CaCO}_3$$

$$\% \text{CaCO}_3 = \frac{2.847 \text{ g}}{3.00 \text{ g}} \times 100\% = \boxed{94.9\%}$$

60 Determine the volume of air (21.0% O$_2$ by volume) needed to burn 227 g of ethanol, C$_2$H$_5$OH, at 35.0°C and 790 mmHg. Write and balance the chemical equation first.

$$\text{C}_2\text{H}_5\text{OH (l)} + 3 \text{O}_2 (g) \rightarrow 2 \text{CO}_2 (g) + 3 \text{H}_2\text{O (l)}$$

$$\text{mol O}_2 = 227 \text{ g ethanol} \times \frac{1 \text{ mol ethanol}}{46.07 \text{ g ethanol}} \times \frac{3 \text{ mol O}_2}{1 \text{ mol ethanol}} = 14.78 \text{ mol O}_2$$

$$V_{O_2} = \frac{nRT}{P} = \frac{(14.78 \text{ mol})(62.36 \frac{\text{L mmHg}}{\text{mol K}})(308 \text{ K})}{790 \text{ mmHg}} = 359.4 \text{ L O}_2$$

$$V_{\text{air}} = 359.4 \text{ L O}_2 \times \frac{100\% \text{ air}}{21.0\% \text{ O}_2} = 1710 \text{ L air} = 1.71 \times 10^3 \text{ L air}$$

130 One oxide of nitrogen has a density of 1.33 g/L at 764 mm Hg and 150°C. What is its formula?

$$M = \frac{dRT}{P} = \frac{(1.33 \frac{g}{L})(0.08206 \frac{\text{L atm}}{\text{mol K}})(150 + 273)K}{764 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 45.9 \text{ g/mol}$$

Some nitrogen oxides and their molar masses are: NO, 30 g/mol; N$_2$O, 44 g/mol; NO$_2$, 46 g/mol

The nitrogen oxide is most likely NO$_2$, although N$_2$O cannot be completely ruled out.