

## Chapter 2: Gasses

1. The atmospheric pressure of **768.2 mm Hg**. Expressed in kilopascals (kPa) what would the value be the pressure?

$$(1 \text{ atm} = 101325 \text{ Pa} = 760 \text{ torr} = 760 \text{ mm Hg})$$

- a. 778.4 kPa
- b. 102.4 kPa**
- c. 100.3 kPa
- d. 91.62 kPa
- e. 1024 kPa

$$\begin{array}{rcl} 101.325 \text{ kPa} & = & 760 \text{ mm Hg} \\ ? & & 768.2 \text{ mm Hg} \\ & & P = 102.4 \text{ kPa} \end{array}$$

2. A sample of a gas occupied a volume of **6.414 liters** when the pressure was **850 torr** and the temperature was  $27.2^\circ\text{C}$ . The pressure was readjusted to **4423 torr**. What was the new **volume**?

- a. 0.837 L
- b. 0.937 L
- c. 1.23 L**
- d. 1.53 L
- e. 3.34 L

$$\begin{array}{ll} V_1=6.414 \text{ L} & P_1= 850 \text{ torr} \\ V_2=? \text{ L} & P_2= 4423 \text{ torr} \end{array}$$

$$\begin{array}{l} P_1.V_1 = P_2.V_2 \\ (850 \text{ torr}) (6.414 \text{ L}) = (V_2) (4423 \text{ torr}) \\ V_2 = 1.23 \text{ L} \end{array}$$

3. A sample of a gas **1.40 liters** when the pressure was **762 torr** and the temperature was  $26.9^\circ\text{C}$ . The volume of the system was readjusted to **0.150 liters**. What was the new **pressure**?

- a. 13.4 atm
- b. 883 atm
- c. 918 atm
- d. 1020 atm
- e. 9.36 atm**

$$\begin{array}{l} P_1.V_1 = P_2.V_2 \\ (762 \text{ torr}) (1.40 \text{ L}) = (x) (0.150 \text{ L}) \\ P = 7112 \text{ torr} \\ P = 7112 / 760 = 9.36 \text{ atm} \end{array}$$

4. A sample of a gas occupied a volume of **1.40 liters** when the pressure was 768 torr and the temperature was **26.9 °C**. The volume of the system was readjusted to **2.16 liters**. What was the **temperature** in the system at this point?

- a. 41.5 °C
- b. 41.9 °C
- c. 189.8 °C**
- d. 194.7 °C
- e. 288.6 °C

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

$$T_2 = V_2 T_1 / V_1$$

$$2.16 \times (26.9 + 273.15) / 1.40 = 462.9 \text{ K}$$

$$462.9 - 273.15 = 189.8 \text{ °C}$$

5. **STP** for gases has the values;

- a. temperature: 0.00 K; pressure: 1.000 standard atmosphere
- b. temperature: 0.00 °C; pressure: 1.000 standard atmosphere**
- c. temperature: 273.15 K; pressure: 1.000 Pascal
- d. temperature: 298.15 K; pressure: 1.000 standard atmosphere
- e. temperature: 298.15 K; pressure: 1.000 Pascal

6. The volume of gas was **1.524 liters** at **28.40 °C**, and **637.6 torr**. What volume would this gas sample occupy at **STP**?

- a. 1.069 L
- b. 1.158 L**
- c. 1.412 L
- d. 1.645 L
- e. 2.006 L

$$\text{STP : } P_2 = 760 \text{ torr and } T_2 = 273 \text{ K}$$

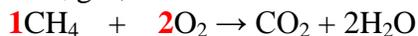
$$P_1 V_1 / T_1 = P_2 V_2 / T_2$$

$$V_2 = (P_1 V_1 T_2) / (P_2 T_1)$$

$$V_2 = 637.6 \times 1.524 \times 273 / 760 \times 301.55$$

$$V_2 = 1.158 \text{ L}$$

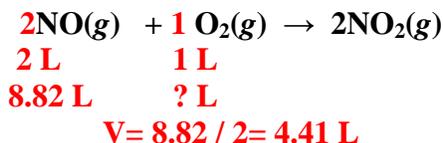
7. How many **liters** of pure **oxygen** gas, measured at STP, are required for the complete combustion of **11.2 L** of **CH<sub>4</sub>** gas, also measured at STP?



$$V = 11.2 \times 2 / 1 = 22.4 \text{ L}$$

- a. 11.2 L
- b. 16.8 L
- c. 22.4 L**
- d. 32.0 L
- e. 33.6 L

8. A chemical reaction is shown:  $2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$ . How many **liters** of pure **oxygen** gas, measured at STP, are required for the complete reaction with **8.82 L** of **NO(g)**, also measured at STP?



- a. **4.41 L**
- b. 8.82 L
- c. 11.2 L
- d. 17.6 L
- e. 22.4 L

9. A gas sample weighing **3.78 grams** occupies a volume of **2.28 L** at **STP**. What is the molecular mass of the sample?

$$\begin{array}{l} P.V = n.R.T \\ PV = mRT/MM \\ MM = mRT/PV \\ MM = [(3.78\text{ g}) (0.08206\text{ L atm mol}^{-1}\text{ K}^{-1}) (273)] / [1\text{ atm}) (2.28\text{ L})] \\ MM = 30.6\text{ g mol}^{-1} \end{array}$$

- a. 8.54 g mol<sup>-1</sup>
- b. 13.5 g mol<sup>-1</sup>
- c. **37.1 g mol<sup>-1</sup>**
- d. 51.1 g mol<sup>-1</sup>
- e. 193 g mol<sup>-1</sup>

10. Two moles of CO<sub>2</sub> gas at 35°C are heated to 250°C, The density of the gas in the gas will:

$$P.MM = d.R.T$$

**The Density is inversely proportional to Temperature**  
**T increases, Density decreases**

- a. increase.
- b. **decrease.**
- c. remain the same.
- d. There is not enough information given to correctly answer this question.

11. What **volume** would **11.2 g** of a gaseous compound occupy at **STP** if its molecular weight is **44.0 g/mole**?

$$\begin{array}{l} P.V = n.R.T \\ PV = mRT/MM \\ V = mRT/PMM \\ V = [(11.2\text{ g}) (0.08206\text{ L atm mol}^{-1}\text{ K}^{-1}) (273)] / [1\text{ atm}) (44.0\text{ g/mole})] \\ V = 5.71\text{L} \end{array}$$

a. **5.71 liters**

b. 11.0 liters

c. 11.2 liters

d. 22.4 liters

e. 44.0 liters

12. A gas sample occupies a volume of **1.66 L** when the temperature is **150.0 °C** and the pressure is **842 torr**. How many molecules are in the sample?

$$P.V = n.R.T$$

$$n = p.V/RT$$

$$n = (842/760)\text{atm} \times 1.66 \text{ L} / 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} (150+273.15 \text{ K})$$

$$n = 0.053 \text{ mol}$$

$$N = n \times N_A = 0.053 \times 6.022 \times 10^{23} = 3.19 \times 10^{22} \text{ molecules}$$

a.  $1.52 \times 10^{22}$

b.  $2.60 \times 10^{22}$

c.  **$3.19 \times 10^{22}$**

d.  $9.01 \times 10^{22}$

e.  $9.42 \times 10^{21}$

13. A gas container has a volume of **6.504 L**. When filled with **C<sub>3</sub>H<sub>8</sub>**, at **28.3 °C**, the pressure is **486.3 torr**. How much should the gas sample weigh?

$$P.V = n.R.T$$

$$P V = m R T / MM$$

$$m = P.V. MM / R T$$

$$m = (486.3/760)\text{atm} \times 6.504 \text{ L} \times (44 \text{ g/mol}) / 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} (28.3 + 273.15 \text{ K})$$

$$m = 7.41 \text{ g}$$

a. 4.67 g

b. **7.41 g**

c. 7.52 g

d. 18.1 g

e. 263. G

14. A container contains **0.2 moles of O<sub>2</sub>** gas and **0.3 moles of N<sub>2</sub>** gas. If the total pressure is **0.75 atm** what is the partial pressure of O<sub>2</sub>?

$$P_i = X_i \cdot P_t$$

$$P_i = (n_i / n_T) \cdot P_t$$

$$P_i = (0.2 / 0.5) 0.75$$

$$P_i = 0.30 \text{ atm}$$

a. 0.20 atm

b. **0.30 atm**

c. 0.50 atm

d. 0.75 atm

e. 0.45 atm

15. A container contains partial pressures of **0.80 atm CO<sub>2</sub>** gas and **0.35 atm N<sub>2</sub>** gas. What is the **mole fraction of N<sub>2</sub>** in the glass container?

$$\begin{aligned} X_i &= P_i / P_t \\ X_i &= 0.35 / (0.8 + 0.35) \\ X_i &= 0.30 \end{aligned}$$

- a. 0.35
- b. 1.15
- c. 0.70
- d. 0.80
- e. 0.30**

16. A gaseous substance **diffuses twice** as rapidly as SO<sub>2</sub> gas. The gas could be

- a. CO
- b. He
- c. H<sub>2</sub>
- d. CH<sub>4</sub>**
- e. O<sub>2</sub>

$$\begin{aligned} \frac{r_1}{r_2} &= \sqrt{\frac{MM_2}{MM_1}} \\ 2 &= \sqrt{\frac{32 + 32}{MM_1}} \\ 4 &= \frac{64}{MM_1} \\ MM_1 &= \frac{64}{4} = 16 \text{ g/mol} \end{aligned}$$

17. According to the kinetic theory of gases, the average **kinetic energy** of the gas particles in a gas sample **is directly proportional** to the;

- a. pressure.
- b. volume.
- c. absolute temperature.**
- d. molar mass.
- e. number of moles of gas.

18. The van der Waals equation of state for a real gas is:  $\left[ P + \frac{n^2 a}{V^2} \right] \left[ \frac{V - nb}{1} \right] = nRT$

At what pressure will 1.00 mole of CH<sub>4</sub> be in a 10.0 L container at 298 K assuming CH<sub>4</sub> is a **real gas**.

(van der Waals constants for CH<sub>4</sub> are  $a = 2.253 \text{ L}^2 \text{ atm mol}^{-2}$ ,  $b = 0.04278 \text{ L mol}^{-1}$ )

- a. 2.43 atm
- b. 2.28 atm
- c. 2.51 atm
- d. 24.5 atm
- e. 0.440 atm

19. A **real gas behaves** most nearly **like an ideal gas** under conditions of
- a. low temperature and high pressure.
  - b. low temperature and low pressure.
  - c. high temperature and low pressure.**
  - d. high temperature and high pressure.
  - e. Actually it will behave like an ideal gas regardless of the temperature or the pressure as long as it remains in the gaseous state.