

Chapter 7

Problems

Gases are used in several medical applications.

Nitrous oxide (N₂O), sold under the brand name **Entonox** among others, is an inhaled gas used as a **pain medication** and together with other **medications** for **anesthesia**. In popular terms, it is called “laughing gas”. (Wikipedia)

Shooting a quick blast of **carbon dioxide** (CO₂) gas into the nose may ease some **allergy symptoms**, and the relief appears to last for about four hours. (Web Md)

Heliox is a **breathing gas** mixture of **helium** (He) and **oxygen** (O₂). **Heliox** is a medical treatment for patients with difficulty breathing. The mixture generates less resistance than atmospheric air when passing through the airways of the lungs, and thus requires less effort by a patient to breathe in and out of the lungs. (Wikipedia)

Problem 7.1

Which of the following gases has the highest molar volume at STP? Molar volume is the volume occupied by exactly one mole of the gas. STP stands for standard temperature (0°C) and pressure (1 atm). **Assume ideal behavior. Explicit calculations are not needed!**

- (a). He
- (b) O₂
- (c). N₂O
- (d) CO₂
- *(e). The molar volume is the same for all the gases

Answer

For an ideal gas,

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

Thus, for $n=1$, at a common temperature (0°C) and pressure (1 atm), the volume is the same! If you wished to calculate it, assuming 3 sf,

$$V = \frac{nRT}{P} = \frac{(1.00 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273.15 \text{ K})}{(1.00 \text{ atm})} = 22.4 \text{ L}$$

This is a useful number to remember for calculations: The volume of 1 mole of an ideal gas (i.e., the molar volume) is 22.4 L at STP.

Problem 7.2

Which of the following gases has the **lowest density** at STP? STP stands for standard temperature (0°C) and pressure (1 atm). **Assume ideal behavior. Explicit calculations are not needed!**

- *(a). He
- (b). O₂
- (c). N₂O
- (d). CO₂
- (e). The density is the same for all the gases at STP

Answer

We use the equation introduced in Chapter 7,

$$M = \frac{mRT}{PV}$$

This rearranges, recognizing $d = m/V$

$$d = \frac{PM}{RT}$$

In these equations, M represents the molar mass (units g/mol), m the mass (units g).

Thus, at a common pressure and temperature, the lower the molar mass (M), lower the density (d).

Problem 7.3

The density of nitrous oxide (N₂O) at STP is _____ g/L. Report your answer to 3sf.

Answer

$$d = \frac{PM}{RT}$$

Noting the molar mass of N₂O to be 44.013 g/mol ,

$$d = \frac{(1.00 \text{ atm})(44.013 \frac{g}{mol})}{(0.08206 \frac{L \cdot atm}{mol \cdot K})(273.15 \text{ K})} = 1.96 \frac{g}{L}$$

We can also calculate the density by noting that 1 mole would occupy a volume of 22.4 L at STP giving a density of $44.013 \text{ g} / 22.4 \text{ L} = 1.96 \text{ g/L}$.

Problem 7.4

Which of the following gases has the highest average speed at STP? STP stands for standard temperature (0°C) and pressure (1 atm). **Assume ideal behavior. Explicit calculations are not needed!**

- *(a). He
- (b). O₂
- (c). N₂O
- (d). CO₂
- (e). The average speed is the same for all the gases at STP

Answer

The lighter a molecule, the higher is the average speed at any temperature. Without calculating molecular masses, we note that He is the lightest species in this set.

Problem 7.5

Helium balloons are often used for parties. For a party, 25 balloons need to be filled each with a volume of 1.00 gallon. The pressure and temperature inside the balloons (1.00 atm, 72°F) matches the ambient pressure and temperature. A vendor carries a 0.500 gallon cylinder. What is the pressure of He in the cylinder at 72°F so that all the balloons can be filled? Assume zero leakage!

Answer

Assume that all the helium is transferred from the cylinder to the 25 balloons, whose combined volume is 25.0 gallons.

$$\frac{P_{\text{cylinder}} V_{\text{cylinder}}}{T_{\text{cylinder}}} = \frac{P_{\text{balloons}} V_{\text{balloons}}}{T_{\text{balloons}}}$$

Noting a common temperature,

$$P_{\text{cylinder}} V_{\text{cylinder}} = P_{\text{balloons}} V_{\text{balloons}}$$

This rearranges to

$$P_{\text{cylinder}} = \frac{P_{\text{balloons}} V_{\text{balloons}}}{V_{\text{cylinder}}}$$

Thus,

$$P_{\text{cylinder}} = \frac{(1.00 \text{ atm})(25.0 \text{ gallon})}{(0.500 \text{ gallon})} = 50.0 \text{ atm}$$

Problem 7.6

Oxygen therapy, also known as **supplemental oxygen**, is the use of **oxygen** as a **medical treatment**. This can include for **low blood oxygen**, **carbon monoxide toxicity**, **cluster headaches**, and to maintain enough oxygen while **inhaled anesthetics** are given. (Wikipedia)

A “Size A” cylinder with a volume of 113 L contains O₂ at 27°C and 100.0 atm. The mass of oxygen in this cylinder, in *kg*, is _____. Assume an ideal gas. Report your answer to 3 sf.

Answer

We use the equation

$$PV = nRT$$

A calculation of the number of moles (*n*) will in turn give the mass.

Rearranging the equation above for *n*,

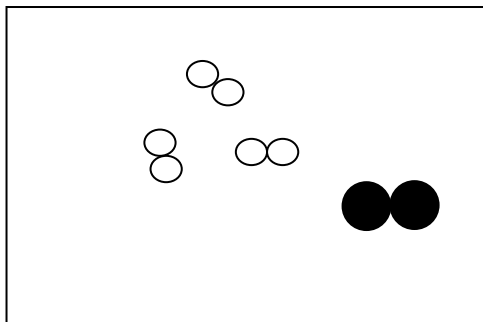
$$n = \frac{PV}{RT}$$

$$n = \frac{(100.0 \text{ atm})(113 \text{ L})}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273.15 + 27)\text{K}} = 458.8 \text{ mol} \quad (\text{Keep an extra digit})$$

$$\text{mass} = 458.8 \text{ mol} \left(\frac{32.00 \text{ g}}{1 \text{ mol}} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = 14.7 \text{ kg} \quad (3 \text{ sf})$$

Problem 7.7

The following sample contains F_2 (open circles) and Br_2 (filled circles). What is the mole fraction of F_2 ?

**Answer**

The sample has one molecule of Br_2 and three molecules of F_2 . Thus, the mole fraction of F_2 is $\frac{3}{4}$.

Problem 7.8

https://www.emedicinehealth.com/decompression_syndromes_the_bends/article_em.htm

(From Wikipedia)

Decompression sickness (DCS; also known as **divers' disease, the bends, aerobullosis, or caisson disease**) describes a condition arising from dissolved gases coming out of solution into bubbles inside the body on depressurization. DCS most commonly refers to problems arising from underwater **diving decompression** (i.e., during ascent), but may be experienced in other depressurization events such as emerging from a **caisson**, flying in an **unpressurised aircraft** at high altitude, and **extravehicular activity** from spacecraft. DCS and **arterial gas embolism** are collectively referred to as **decompression illness**.

Mixtures of helium (He) and oxygen (O₂) are used by divers. For which of the following mixtures is the mole fraction of helium 0.50?

(a) 4.00 g of He and 4.00 g of O₂

(b) 4.00 g of He and 8.00 g of O₂

(c) 4.00 g of He and 16.0 g of O₂

*(d) 4.00 g of He and 32.0 g of O₂

(e) 4.00 g of He and 64.0 g of O₂

Answer

$$\chi_{\text{He}} = 0.50$$

$$\chi_{\text{He}} + \chi_{\text{O}_2} = 1$$

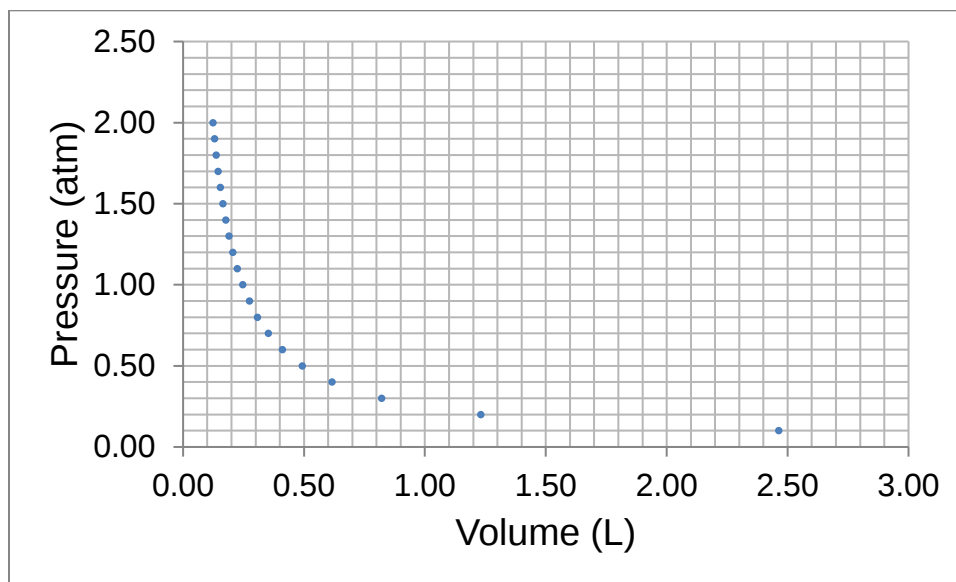
Thus,

$$\chi_{\text{He}} = \chi_{\text{O}_2} = 0.50$$

Another way to approach this is to see which sample has equal number of moles for the two components. The molar mass of He (to 3 sf) is 4.00 g/mol and that for O₂ is 32.0 g/mol. Thus, in choice (d), we note an equal number of moles for each species.

Problem 7.9

A 0.320 g sample of O₂ was placed in a cylinder connected to a manometer (a device that measures pressure). The volume of the container can be varied using a piston. The following data were obtained, while the temperature was kept constant.



- (a). This experiment illustrates _____ Law.
(b). The temperature in this experiment is _____ K. Estimate to 2 sf.

Answer

(a). These data are acquired at a constant temperature for a constant mass of gas. As the pressure decreases, the volume increases. Thus, this experiment illustrates Boyle's Law.

(b). We can pick any point. We note that when the volume is 0.50 L, the pressure is *very close* to 0.50 atm.

We use the equation

$$PV = nRT$$

Rearranging the equation above for T , and noting that 0.320 g O₂ corresponds to 0.0100 mol O₂,

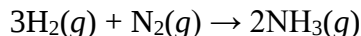
$$T = \frac{PV}{nR}$$

$$T = \frac{(0.50 \text{ atm})(0.50 \text{ L})}{(0.0100 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})} = 305 \text{ K}$$

We can report this to $3.0 \times 10^2 \text{ K}$ (2sf).

Problem 7.10

As discussed in Lesson 6, the production of ammonia (NH₃) from nitrogen and hydrogen, shown below, was an important breakthrough resulting in Nobel prizes in chemistry to Fritz Haber (1918) and Carl Bosch (1931). This compound is widely used in manufacturing fertilizers.



In an experiment conducted at STP, 10.0 L of N₂ at STP reacts completely with H₂ at STP to produce NH₃. The volume of H₂ needed is ____ L and (assuming 100% yield) the amount of NH₃ produced is ____ L.

Answer

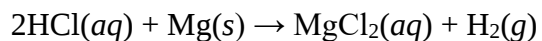
This is a stoichiometry problem. Without calculating moles explicitly, we can carry out the calculations using Avogadro's Law: *For a confined gas, the volume (V) and number of moles (n) are directly proportional if the pressure and temperature both remain constant.*

Thus, 10.0 L of any gas contains the same number of moles! The number of moles of H₂ is three times those of N₂, and the number of moles of NH₃ is twice those of N₂, based on the stoichiometric coefficients in the balanced equation.

10.0 L of N₂ will react with 30.0 L of H₂ to produce 20.0 L of NH₃.

Problem 7.11

A flask contains 0.100 moles of an aqueous solution of hydrochloric acid (HCl). A large ribbon of magnesium is carefully placed in the HCl and the hydrogen gas thus generated is trapped in a balloon. Calculate the volume of hydrogen generated at STP, given the balanced equation below. Assume that all the HCl is consumed.

**Answer**

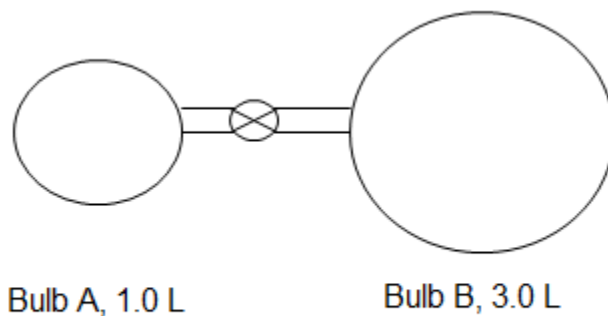
This is a stoichiometry problem.

$$n_{\text{H}_2} = (0.100 \text{ mol HCl}) \left(\frac{1 \text{ mol H}_2}{2 \text{ mol HCl}} \right) = 0.0500 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(0.0500 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273.15 \text{ K})}{(1.00 \text{ atm})} = 1.12 \text{ L}$$

Problem 7.12

Two glass bulbs are shown below.



Initially, the pressure of N_2 in Bulb A is 1.00 atm, and Bulb B is evacuated (i.e., has no gas). The valve between bulbs A and B is **closed**. What is the **final** pressure in the two bulbs when the valve is **opened**? Assume ideal behavior, and a constant temperature. Ignore the volume of the valve and the tubing connecting the two bulbs.

Answer

This is a “disguised” problem on Boyle’s Law!

Noting a common temperature,

$$P_{initial} V_{initial} = P_{final} V_{final}$$

This rearranges to

$$P_{final} = \frac{P_{initial} V_{initial}}{V_{final}}$$

$$P_{final} = \frac{(1.00 \text{ atm})(1.0 \text{ L})}{(4.0 \text{ L})} = 0.25 \text{ atm} \quad (2 \text{ sf})$$