

This print-out should have 20 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

Nlib 04 0022

001 10.0 points

Two gases are contained in gas bulbs connected by a valve. Gas *A* is present in a 1 liter bulb at a pressure of 811 torr. Gas *B* exerts a pressure of 277 torr in a 1 liter bulb. The valve is opened and the two gases equilibrate. What is the partial pressure of gas *A* expressed after equilibration?

Correct answer: 405.5 torr.

Explanation:

$$V_A = 1 \text{ L}$$

$$P_A = 811 \text{ torr}$$

$$V_{\text{total}} = 2 \text{ L}$$

$$V_B = 1 \text{ L}$$

$$P_B = 277 \text{ torr}$$

$$\begin{aligned} P_A V_A &= P_t V_t \\ P_t &= \frac{P_A V_A}{V_t} \\ &= \frac{(811 \text{ torr})(1 \text{ L})}{2 \text{ L}} = 405.5 \text{ torr} \end{aligned}$$

ChemPrin3e T04 49

002 10.0 points

A reminder... partial pressures are calculated from the total pressure via mole fractions, not percent by mass. So you WILL have to convert the percent by mass to mole fraction BEFORE you answer this problem.

A mixture of oxygen and helium is 92.3% by mass oxygen. It is collected at atmospheric pressure (745 torr).

What is the partial pressure of oxygen in this mixture?

1. 412 Torr

2. 448 Torr **correct**

3. 688 Torr

4. 333 Torr

5. 299 Torr

Explanation:

Assume you have 100 g of this mixture; calculate the number of moles:

$$\begin{aligned} n_{\text{O}_2} &= (92.3 \text{ g O}_2) \frac{1 \text{ mol O}_2}{31.9988 \text{ g O}_2} \\ &= 2.91261 \text{ mol O}_2. \end{aligned}$$

$$\begin{aligned} n_{\text{He}} &= (7.7 \text{ g He}) \frac{1 \text{ mol He}}{4.0026 \text{ g He}} \\ &= 1.92375 \text{ mol He.} \end{aligned}$$

$$\begin{aligned} n_{\text{tot}} &= n_{\text{O}_2} + n_{\text{He}} \\ &= 2.91261 \text{ mol O}_2 + 1.92375 \text{ mol He} \\ &= 4.83636 \text{ mol gas} \end{aligned}$$

Dalton's Law:

$$\begin{aligned} P_{\text{O}_2} &= P_{\text{tot}} \times \chi_{\text{O}_2} \\ &= P_{\text{tot}} \times \frac{n_{\text{O}_2}}{n_{\text{tot}}} \\ &= (745 \text{ Torr}) \frac{2.91261 \text{ mol O}_2}{4.83636 \text{ mol gas}} \\ &= 448.663 \text{ Torr} \end{aligned}$$

ChemPrin3e T04 57

003 10.0 points

If the average speed of a water molecule at 25°C is 640 m · s⁻¹, what is the average speed at 100°C?

1. 716 m/s **correct**

2. 801 m/s

3. 572 m/s

4. 320 m/s

5. 5120 m/s

Explanation:

$$T_1 = 100^\circ\text{C} + 273.15 = 373.15 \text{ K}$$

$$T_2 = 25^\circ\text{C} + 273.15 = 298.15 \text{ K}$$

From kinetic molecular theory, the temperature is *directly* proportional to mean KE.

$\text{KE}_{\text{mean}} = \frac{1}{2}(\text{MW})(\text{average molecular speed})^2$
and knowing MW is constant (it's the same gas) we get average speed v_T of gas molecules:

$$\frac{v_{T_1}}{v_{T_2}} = \frac{\sqrt{T_1}}{\sqrt{T_2}}$$

$$v_{T_2} = v_{T_1} \frac{\sqrt{T_2}}{\sqrt{T_1}} = (640 \text{ m/s}) \sqrt{\frac{373.15 \text{ K}}{298.15 \text{ K}}} \\ = 715.986 \text{ m/s}$$

ChemPrin3e 04 90

004 (part 1 of 2) 10.0 points

Consider (calculate) ONLY the nitrogen gas that comes from the azide in this problem.

Air bags in automobiles contain crystals of sodium azide (NaN_3) which during a collision decompose rapidly to give nitrogen gas and sodium metal. Potassium nitrate and silicon dioxide are added to remove the sodium metal by converting it into a harmless material. First, the sodium reacts with potassium nitrate to produce potassium oxide (K_2O), sodium oxide (Na_2O), and additional nitrogen gas. The metal oxides (K_2O and Na_2O) react with silicon dioxide in a final reaction to produce silicate glass, which is harmless and stable. The nitrogen gas liberated by this process instantly inflates the air bag. Assume that the nitrogen gas liberated behaves as an ideal gas and that any solid produced has a negligible volume (which may be ignored). Calculate the mass of sodium azide required to generate enough nitrogen gas to fill a 46 L air bag at 1.45 atm and 39°C .

Correct answer: 112.926 g.

Explanation:

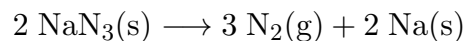
$$P = 1.45 \text{ atm} \quad T = 39^\circ\text{C} + 273 = 312 \text{ K} \\ R = 0.08206 \text{ L} \cdot \text{atm}/\text{K}/\text{mol} \quad V_{\text{N}_2} = 46 \text{ L} \\ \text{MW} = 65.02 \text{ g}$$

Applying the ideal gas law,

$$n_{\text{N}_2} = \frac{PV}{RT}$$

$$= \frac{(1.45 \text{ atm})(46 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm}/\text{K}/\text{mol})(312 \text{ K})} \\ = 2.60519 \text{ mol N}_2.$$

The balanced reaction is



$$m_{\text{NaN}_3} = (2.60519 \text{ mol N}_2) \left(\frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \right) \\ \times \left(\frac{65.02 \text{ g NaN}_3}{1 \text{ mol NaN}_3} \right) \\ = 112.926 \text{ g NaN}_3.$$

005 (part 2 of 2) 10.0 points

What is the root mean square speed of the N_2 gas molecules generated?

Correct answer: 527.186 m/s.

Explanation:

$$\text{MW}_{\text{N}_2} = 28 \text{ g/mol} \quad R = 8.314 \text{ J}/\text{K}/\text{mol}$$

$$V_{rms} = \sqrt{\frac{3RT}{\text{MW}_{\text{N}_2}}} \\ = \sqrt{\frac{3(8.314 \text{ J}/\text{K}/\text{mol})(312 \text{ K})}{(28 \text{ g/mol})}} \\ \times \sqrt{\left(\frac{\text{kg}/\text{m}^2/\text{s}^2}{\text{J}} \right) \left(\frac{1000 \text{ g}}{\text{kg}} \right)} \\ = 527.186 \text{ m/s}.$$

Mlib 04 1183

006 10.0 points

Which of the following statements about the Kinetic-Molecular Theory of gases is false?

1. The average kinetic energy of a gas molecule is independent of the temperature.
correct
2. Collisions between molecules are elastic.

3. Gases consist of molecules in continuous, random motion.

4. The distance between molecules is much larger than the diameter of each molecule.

Explanation:

The average kinetic energy of gas molecules is very much dependent on temperature. The higher the temperature, the larger is the average kinetic energy of the gas molecules.

ChemPrin3e T04 62

007 10.0 points

A plot of the Maxwell distribution for the same gas against temperature shows that

1. at low temperatures, most molecules have speeds close to their average speed. **correct**

2. at high temperatures, most molecules have speeds close to their average speed.

3. as the temperature decreases, the spread of speeds widens.

4. as the temperature decreases, a high proportion of molecules have very high speeds.

5. as the temperature increases, a high proportion of molecules have very slow speeds.

Explanation:

The Maxwell distribution of speeds for a particular sample of gas at a given temperature shows the following characteristics if one of the following variables is changed:

Mass: The *distribution* of speeds for a given sample of gas is *narrower* and the *average* speed is *lower* for a gas of *higher* molecular weight.

Temperature: The *distribution* of speeds for a given sample of gas is *wider* and the *average* speed is *higher* as the temperature of a gas *increases*.

Msci 12 1306 ran

008 10.0 points

Given the gases Ne, N₂, Cl₂, CO₂ put them

in order of their DECREASING rate of effusion.

1. N₂ > Ne > Cl₂ > CO₂

2. Ne > N₂ > CO₂ > Cl₂ **correct**

3. CO₂ > Cl₂ > Ne > N₂

4. Cl₂ > CO₂ > N₂ > Ne

Explanation:

Effusion is when gas molecules escape through the tiny holes in their container. Lighter, faster molecules effuse more quickly than heavy, slow ones.

$$\begin{aligned} \text{Ne}_{(\text{MW}=20 \text{ g/mol})} &> \text{N}_{2(\text{MW}=28 \text{ g/mol})} \\ &> \text{CO}_{2(\text{MW}=44 \text{ g/mol})} \\ &> \text{Cl}_{2(\text{MW}=71 \text{ g/mol})} \end{aligned}$$

Msci 12 1305

009 10.0 points

Calculate the ratio of the rate of effusion of CO₂ to that of He.

1. 0.090 : 1

2. 12 : 1

3. 3.3 : 1

4. 11 : 1

5. 0.30 : 1 **correct**

Explanation:

$$\frac{\text{Eff}_{\text{CO}_2}}{\text{Eff}_{\text{He}}} = \frac{\sqrt{\text{MW}_{\text{He}}}}{\sqrt{\text{MW}_{\text{CO}_2}}} = \sqrt{\frac{4}{44}} = \frac{0.30}{1}$$

DVB Kinetic Theory of Gases 001

010 10.0 points

A sample of He gas and O₂ have the same temperature, pressure, and volume. Which gas has a greater number of collision of gas molecules with the walls of the container?

1. The O₂ since it has a higher average kinetic energy because it is more massive

2. The He gas because it is less massive and moving with a higher average velocity **correct**

3. The O₂ since it has a higher average momentum since it is more massive

4. They are the same since the pressure is the same

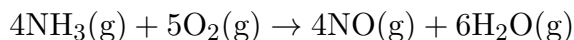
Explanation:

Both gases have the same temperature so both have the same average kinetic energy. The lighter mass of He means it will have a higher velocity. The higher velocity will mean the He gas will have more collisions. However, the net result will be the same pressure as each collision will have less force as the molecules will have a smaller momentum.

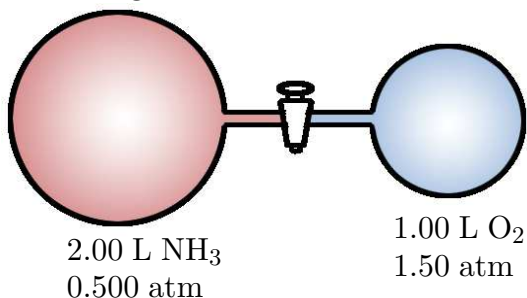
Ostwald process 01

011 10.0 points

Nitric acid is produced commercially by the Ostwald process. In the first step, ammonia is oxidized to nitric oxide:



Assume this reaction is carried out in the apparatus diagramed below.



The stopcock between the two reaction containers is opened, and the reaction proceeds using proper catalysts. Calculate the partial pressure of NO after the reaction is complete. Assume 100% yield for the reaction, assume the final container volume is 3.00 L, and assume the temperature is constant.

1. 1.50 atm

2. 0.250 atm

3. 0.500 atm

4. 0.400 atm

5. 0.333 atm **correct**

Explanation:

First, calculate the partial pressures of NH₃ and O₂ after the stopcock is opened (and before the reaction). Use Boyle's Law... new volume is 3 L so that...

$$P_{\text{NH}_3} = (2\text{L}/3\text{L})(1/2\text{atm}) = 1/3 \text{ atm}$$

$$P_{\text{O}_2} = (1\text{L}/3\text{L})(3/2\text{atm}) = 1/2 \text{ atm}$$

You must have at least 5/4 more moles of O₂ than NH₃ for complete reaction. That means you need 5/12 atm of O₂. You have more than enough (0.5 atm = 6/12 atm) so the oxygen is in excess.

The easy part... You make the SAME number of moles of NO as you use NH₃, so the final partial pressure of the NO will be the same as the initial partial pressure of the NH₃ which is 1/3 atm or 0.333 atm.

Brodbeck 013 511

012 10.0 points

A 5.0 L flask containing O₂ at 2.00 atm is connected to a 3.0 L flask containing H₂ at 4.00 atm and the gases are allowed to mix. What is the mole fraction of H₂?

1. 0.67

2. 0.55 **correct**

3. 0.25

4. 0.45

5. 0.33

6. Cannot be determined

Explanation:

Applying the ideal gas law

$$\begin{aligned}
 PV &= nRT \\
 &= \frac{PV}{RT}
 \end{aligned}$$

$$N_{\text{O}_2} = \frac{(2 \text{ atm})(5 \text{ L})}{RT}$$

$$N_{\text{H}_2} = \frac{(4 \text{ atm})(3 \text{ L})}{RT}$$

Assume the temperature of the two gases remains the same before and after the mixing occurs. The mole fraction of H₂ is

$$\chi_{\text{H}_2} = \frac{n_{\text{H}_2}}{n_{\text{total}}} = \frac{n_{\text{H}_2}}{n_{\text{H}_2} + n_{\text{O}_2}}$$

$$= \frac{\frac{(4 \text{ atm})(3 \text{ L})}{RT}}{\frac{(4 \text{ atm})(3 \text{ L})}{RT} + \frac{(2 \text{ atm})(5 \text{ L})}{RT}}$$

$$= \frac{(4 \text{ atm})(3 \text{ L})}{(4 \text{ atm})(3 \text{ L}) + (2 \text{ atm})(5 \text{ L})}$$

$$= 0.545455$$

ChemPrin3e 04 56

013 (part 1 of 2) 10.0 points

A gas mixture being used to simulate the atmosphere of another planet consists of 359 mg of methane, 176 mg of argon, and 241 mg of nitrogen. The partial pressure of nitrogen at 297 K is 10 kPa. Calculate the total pressure of the mixture.

Correct answer: 41.133 kPa.

Explanation:

$$m_{\text{N}_2} = 241 \text{ mg} = 0.241 \text{ g} \quad P_{\text{N}_2} = 10 \text{ kPa}$$

$$m_{\text{CH}_4} = 359 \text{ mg} = 0.359 \text{ g}$$

$$m_{\text{Ar}} = 176 \text{ mg} = 0.176 \text{ g}$$

$$n_{\text{CH}_4} = \frac{0.359 \text{ g}}{16.04 \text{ g/mol}} = 0.0223815 \text{ mol}$$

$$n_{\text{Ar}} = \frac{0.176 \text{ g}}{39.95 \text{ g/mol}} = 0.00440551 \text{ mol}$$

$$n_{\text{N}_2} = \frac{0.241 \text{ g}}{28.01 \text{ g/mol}} = 0.00860407 \text{ mol}$$

$PV = nRT$, so $P \propto n$ and the partial pressures are

$$P_{\text{N}_2} = 10 \text{ kPa}$$

$$P_{\text{CH}_4} = \frac{n_{\text{CH}_4}}{n_{\text{N}_2}} \times P_{\text{N}_2}$$

$$= \frac{0.0223815 \text{ mol}}{0.00860407 \text{ mol}} \times 10 \text{ kPa}$$

$$= 26.0127 \text{ kPa}$$

$$P_{\text{Ar}} = \frac{n_{\text{Ar}}}{n_{\text{N}_2}} \times P_{\text{N}_2}$$

$$= \frac{0.00440551 \text{ mol}}{0.00860407 \text{ mol}} \times 10 \text{ kPa}$$

$$= 5.12026 \text{ kPa}$$

$$P_{\text{total}} = 26.0127 \text{ kPa} + 5.12026 \text{ kPa}$$

$$+ 10 \text{ kPa}$$

$$= 41.133 \text{ kPa}.$$

014 (part 2 of 2) 10.0 points

Calculate the volume.

Correct answer: 2.12475 L.

Explanation:

$$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \quad T = 297 \text{ K}$$

Using the total moles and pressure,

$$n_{\text{total}} = 0.0223815 \text{ mol} + 0.00440551 \text{ mol}$$

$$+ 0.00860407 \text{ mol}$$

$$= 0.0353911 \text{ mol}.$$

$$PV = nRT$$

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

$$V = \frac{(0.0353911 \text{ mol})(297 \text{ K})}{41.133 \text{ kPa}}$$

$$\times 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$$

$$\times \frac{101.325 \text{ kPa}}{1 \text{ atm}}$$

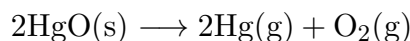
$$V = 2.12475 \text{ L}.$$

Note: You can also obtain the same result using any of the three separate gases and its partial pressure.

DVB Dalton's Law RXN 001

015 10.0 points

When heated, solid mercury oxide HgO will decompose into mercury and oxygen gas



You start with a container that has only HgO solid in it. The temperature is raised to 700K and you find that all of the solid has decomposed (and all the mercury is vapor). The total pressure in the container is 0.75 bar. What is the partial pressure of oxygen??

1. 0.375 bar
2. 0.75 bar
3. 0.25 bar **correct**
4. 0.125 bar
5. 0.5 bar

Explanation:

The number of moles of mercury vapor that is formed is twice the number of moles of O₂. Therefore the partial pressure of Hg(g) will be twice that of O₂. So $P_{Hg} = 2P_{O_2}$. The total pressure is the sum of the partial pressures, $P_{Hg} + P_{O_2} = 2P_{O_2} + P_{O_2} = 3 \times P_{O_2} = 0.75 \text{ bar}$

$$P_{O_2} = 0.25 \text{ bar}$$

DVB Ideal Gas Model 001

016 10.0 points

In the limit of low pressure, all gases exhibit ideal behavior because when the pressure is very low the gas particles are

1. far apart and rarely interacting **correct**
2. undergoing only elastic collisions
3. experiencing a balance of kinetic and potential energy
4. slightly attracted to one another
5. moving very slowly

Explanation:

At very low pressures the average distance between gas particles is large and therefore the particles are rarely interacting with each other. As such, all gases appear to have the same behavior since they are essentially

invisible to one another.

Gases move slower at lower temperatures

Generally all substances are slightly attracted to themselves. This does not vary with pressure. In a real gas attraction are generally important at intermediate pressure as attractive forces are longer range than repulsive forces.

DVB Ideal Gas Model 002

017 10.0 points

Deviations from ideal gas behavior can be modeled with other equations of state. One such equation that attempts to account for the repulsive interactions of gas particles is the hard sphere model

$$P(V - nb) = nRT$$

A 1 mole sample of He gas at 1000K and 500 bar has a volume of 0.176 L. Estimate the value of the constant b in the hard sphere model for He?

1. .01 L mol⁻¹ **correct**
2. .025 L mol⁻¹
3. .002 L mol⁻¹
4. .176 L mol⁻¹
5. .166 L mol⁻¹

Explanation:

Rearranging the hard sphere model gives $V = (nRT/P) + nb$

$$V = V_{IG} + nb$$

That is the volume of the ideal gas + nb

$$nb = V - V_{IG}$$

$$V_{IG} = (1)(.08314)(1000)/(500) = 0.166$$

$$nb = 0.176 - 0.166 = 0.01 \text{ L}$$

$$0.01 \text{ L}/1 \text{ mol} = 0.01 \text{ L mol}^{-1}$$

DVB Ideal Gas Model 003

018 10.0 points

The ideal gas equation models the gas behavior observed in the real world

1. perfectly for some gases but not for oth-

ers

2. perfectly for pressures under 20 atm
3. very well under some conditions but shows large errors in others **correct**
4. perfectly under all conditions for all gases
5. perfectly for temperatures under 1000 K

Explanation:

The ideal gas model is never perfect. In the limit that the pressure goes to zero all gases exhibit ideal behavior. But at finite pressure there are always deviations (albeit very small at low pressures).

ChemPrin3e T04 66

019 10.0 points

Which of the following gases would you predict to have the largest value of the van der Waals coefficient b ?

1. C_2F_6
2. Cl_2
3. CO_2
4. C_2FCl_5 **correct**
5. $C_2F_2Cl_4$

Explanation:

The larger the van der Waals coefficient b , the larger the molecule. The largest molecules have 8 atoms. Of these, C_2FCl_5 is larger than $C_2F_2Cl_4$ and C_2F_6 , because Cl is larger than F.

ChemPrin3e T04 64

020 10.0 points

Consider the following van der Waals coefficients:

Gas	$a, L^2 \cdot atm \cdot mol^{-2}$	$b, L \cdot mol^{-1}$
helium	0.034	0.0237
hydrogen	0.244	0.0266
neon	0.211	0.0171
krypton	2.32	0.0398
xenon	4.19	0.0511
chlorine	6.49	0.0562
CO_2	3.59	0.0427
ammonia	4.17	0.0371
water	5.46	0.0305

Which of the following gases has the largest attractive forces?

1. chlorine **correct**
2. helium
3. neon
4. ammonia
5. water

Explanation:

The larger the Van der Waals coefficient a , the larger the attractive forces.