Chapter 12

Exploring Gas Laws

Solutions for Practice Problems

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1. **Problem**

   At 19°C and 100 kPa, 0.021 mol of oxygen gas, $O_{2(g)}$, occupy a volume of 0.50 L. What is the molar volume of oxygen gas at the temperature and pressure?

   What Is Required?

   You need to find the volume of one mole of oxygen at the given temperature and pressure.

   What Is Given?

   P = 100 kPa
   T = 19°C
   V = 0.50 L
   n = 0.021 mol

   Plan Your Strategy

   Divide the volume (V) by the number of moles (n) to get the molar volume.

   Act on Your Strategy

   $\frac{V}{n} = \frac{0.50 \text{ L}}{0.021 \text{ mol}} = 23.8 \text{ L/mol}$

   Therefore, the molar volume of oxygen gas at the given temperature and pressure is 23.8 L/mol.

   Check Your Solution

   The molar volume is the volume of one mole of gas. The units in the answer are in L/mol, which is consistent with this. A quick round number check of the numbers, 5 / 0.2 = 25, gives a value close to the 23.8 calculated. Your answer is reasonable.

2. **Problem**

   What is the molar volume of hydrogen gas, $H_{2(g)}$, at 255°C and 102 kPa, if a 1.09 L volume of the gas has a mass of 0.0513 g?

   What Is Required?

   You need to find the volume of one mole of hydrogen at the given temperature and pressure.

   What Is Given?

   T = 255°C
   P = 102 kPa
   V = 1.09 L
   m = 0.0513 g

   Plan Your Strategy

   Calculate the number of moles (n) of hydrogen by dividing its given mass (m) by its molar mass (M). The n divide this result into the given volume of the gas for the molar volume.

Act on Your Strategy

$n = \frac{m}{M} = \frac{0.0513 \cancel{g}}{2.02 \cancel{g}/mol} = 0.0254 \text{ mol } H_2$

Therefore, $\frac{V}{n} = \frac{1.09 \text{ L}}{0.0254 \text{ mol}} = 42.9 \text{ L/mol}$

The molar volume of the H_2 at the given temperature and pressure is 42.9 L/mol.

Check Your Solution

The molar volume is the volume of one mole of gas. The units in the answer are in L/mol, which is consistent with this. A quick round number check of the numbers, $1 / 0.025 = 40$, gives a value close to the 42.9 calculated. Your answer is reasonable.

3. Problem

A sample of helium gas, $He_{2(g)}$, has a mass of 11.28 g. At STP, the sample has a volume of 63.2 L. What is the molar volume of this gas at 32.2°C and 98.1 kPa?

What Is Required?

You need to find the volume of one mole of helium at the given temperature and pressure.

What Is Given?

You can set out the data in a table like the one below.

Situation at STP (*i*)	Situation in the container (*f*)
T_i = 25°C or 298 K	T_f = 32.3°C or 305.2 K
P_i = 101.3 kPa	P_f = 98.1 kPa
V_i = 63.2 L	V_f = ?
m = 11.28 g	*m* = 11.28 g

Plan your Strategy

Step 1 First calculate the number of moles (n) by dividing the given mass (m) of helium by its molar mass (M).
Step 2 Apply the combined gas law from Chapter 11 to solve for V_f .
Step 3 Divide the answer in Step 2 by n to obtain the molar volume.

Act on Your Strategy

Step 1 $n = \frac{m}{M} = \frac{11.28 \cancel{g}}{4.00 \cancel{g}/mol} = 2.82 \text{ mol}$

Step 2 $\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$

$V_f = \frac{P_i\ V_i\ T_f}{T_i\ P_f} = \frac{101.3 \cancel{\text{kPa}} \times 63.2 \text{ L} \times 305.2 \cancel{\text{K}}}{298 \cancel{\text{K}} \times 98.1 \cancel{\text{kPa}}} = 66.8 \text{ L}$

Step 3 $\frac{V}{n} = \frac{66.8 \text{ L}}{2.82 \text{ mol}} = 23.7 \text{ L/mol}$

Therefore, the molar volume of helium at the given temperature and pressure is 23.7 L/mol.

Check Your Solution

The molar volume is the volume of one mole of gas. The units in the answer are in L/mol, which is consistent with this. A quick round number check of the numbers in step 3, $\frac{66}{3} = 22$, gives a value close to the 23.7 calculated. Your answer is reasonable.

4. Problem

In the Sample Problem, you discovered that the molar volume of nitrogen gas is 22.4 L at STP.

(a) How many moles of nitrogen are present in 10.0 L at STP?
(b) What is the mass of this gas sample?

What Is Required?

(a) You need to find the number of moles of N_2 at STP.

(b) You need to find the mass of N_2.

What Is Given?

(a) The molar volume calculated in the Sample Problem 22.4 L/mol, i.e., $V_i = 22.4$ L; $n_i = 1$ mol. $V_f = 10.0$ L.

(b) The molar mass of nitrogen = 28.02 g/mol.

Plan Your Strategy

(a) Calculate the number of moles of nitrogen in 10.0 L at STP by using Avogadro's law and solving for n_f .

(b) Multiply n_f by the molar mass of nitrogen to obtain the mass.

Act on Your Strategy

(a) $\frac{n_i}{V_i} = \frac{n_f}{V_i} \therefore n_f = \frac{1 \text{ mol} \times 10.0 \cancel{\text{L}}}{22.4 \cancel{\text{L}}} = 0.446 \text{ mol}$

(b) $m = n_f \times M = 0.446 \cancel{\text{mol}} \times 28.02 \text{ g}/\cancel{\text{mol}} = 12.5 \text{ g of } N_2$.

Check Your Solution

(a) $\frac{V_f}{n_f} = \frac{10.0 \text{ L}}{0.446 \text{ mol}} = 22.4 \text{ L/mol}$. This molar volume at STP is the same as that given in the question. The answer matches.

(b) The mass of 1 mol of N_2 is 28.02 g. The mass of 0.446 mol is calculated to be 12.5 g. Working with a ratio check, $\frac{28.02 \text{ g}}{12.5 \text{ g}} = 2.24$, while $\frac{1 \text{ mol}}{0.446 \text{ mol}} = 2.24$. The ratios match, and your answer is consistent.

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5. Problem

A balloon contains 2.0 L of helium gas at STP. How many moles of helium are present?

What Is Required?

You need to find the number of moles of helium at STP.

What Is Given?

The gas is at STP, so its molar volume = 22.4 L/mol. That is, $V_i = 22.4$ L, $n_i = 1$ mol, $V_f = 2.0$ L.

Plan Your Strategy

Use Avogadro's law to solve for n_f.

Act on Your Strategy

$\frac{n_i}{V_i} = \frac{n_f}{V_i} \therefore n_f = \frac{1 \text{ mol} \times 2.0 \cancel{\text{L}}}{22.4 \cancel{\text{L}}} = 0.089 \text{ mol}$

Therefore, the number of moles of helium is 0.089 mol.

Check Your Solution

The molar volume is 22.4 L/mol. Working with the numbers, $\frac{2.0 \text{ L}}{0.089 \text{ mol}} = 22.4 \text{ L/mol}$. The results match.

6. Problem

How many moles of gas are present in 11.2 L at STP? How many molecules?

What Is Required?

(a) You need to find the number of moles of gas at the given volume at STP.

(b) You need to find the number of molecules of gas at STP.

What Is Given?

(a) At STP, $V_i = 22.4$ L; $n_i = 1$ mol; $V_f = 11.2$ L

(b) You know Avogadro's number = 6.02×10^{23} molecules/mol.

Plan Your Strategy

(a) Use Avogadro's law to solve for n_f.

(b) Multiply n_f by Avogadro's number to obtain the number of molecules at this given volume.

Act on Your Strategy

$$\frac{n_i}{V_i} = \frac{n_f}{V_f} \therefore n_f = \frac{1 \text{ mol} \times 11.2 \cancel{\text{L}}}{22.4 \cancel{\text{L}}} = 0.500 \text{ mol}$$

Number of molecules of gas $= 0.500 \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}}$

$= 3.01 \times 10^{23}$ molecules

Check Your Solution

The molar volume is 22.4 L/mol. Working with the numbers,

$\frac{11.2 \text{ L}}{0.500 \text{ mol}} = 22.4$ L/mol. The results match.

7. **Problem**

What is the volume, at STP, of 3.45 mol of argon gas?

What Is Required?

You need to find the volume of the argon gas for the given number of moles.

What Is Given?

At STP, $V_i = 22.4$ L; $n_i = 1$ mol; $n_f = 3.45$ mol

Plan Your Strategy

Use Avogadro's law to solve for V_f.

Act on Your Strategy

$$\frac{n_i}{V_i} = \frac{n_f}{V_f} \therefore V_f = \frac{3.45 \cancel{\text{mol}} \times 22.4 \text{ L}}{1.00 \cancel{\text{mol}}} = 77.3 \text{ L}$$

Check Your Solution

The molar volume is 22.4 L/mol. Working with the numbers, $\frac{77.3 \text{ L}}{3.45 \text{ mol}} = 22.4$ L/mol. The results match.

8. **Problem**

A certain set of conditions allows 4.0 mol of gas to be held in a 70 L container. What volume do 6.0 mol of gas need under the same conditions of temperature and pressure?

What Is Required?

You need to find the volume of the gas for the given number of moles.

What Is Given?

$V_i = 70$ L

$n_i = 4.0$ mol

$n_f = 6.0$ mol

Plan Your Strategy

Use Avogadro's law to solve for V_f.

Act on Your Strategy

$$\frac{n_i}{V_i} = \frac{n_f}{V_f} \therefore V_f = \frac{6.0 \cancel{\text{mol}} \times 70 \text{ L}}{4.0 \cancel{\text{mol}}} = 105 \text{ L}$$

Check Your Solution

Working with the ratios, $\frac{V_i}{n_i} = \frac{70 \text{ L}}{4.0 \text{ mol}} = 17.5$ L/mol; $\frac{V_f}{n_f} = \frac{105 \text{ L}}{6 \text{ mol}} = 17.5$ L/mol.

Both molar volumes match.

9. Problem

At STP, a container holds 14.01 g of nitrogen gas, 16.00 g of oxygen gas, 66.00 g of carbon dioxide gas, and 17.04 g of ammonia gas. What is the volume of the container?

What Is Required?

You need to find the volume of the container.

What Is Given?

At STP, $V_i = 22.4$ L; $n_i = 1$ mol

The mass of each gas is given. You have access to the periodic table for the molar masses of the elements.

Plan Your Strategy

The volume of the container will comprise the total number of moles of all the gases present. Divide each mass of gas by its own molar mass to obtain its number of moles. Add the number of moles together, and then use this as n_f in Avogadro's law to solve for V_f.

Act on Your Strategy

$$n_{nitrogen} = \frac{14.01\ \cancel{g}}{28.02\ \cancel{g}/mol} = 0.500\ mol$$

$$n_{oxygen} = \frac{16.00\ \cancel{g}}{32.00\ \cancel{g}/mol} = 0.500\ mol$$

$$n_{carbon\ dioxide} = \frac{66.00\ \cancel{g}}{44.01\ \cancel{g}/mol} = 1.50\ mol$$

$$n_{ammonia} = \frac{17.04\ \cancel{g}}{17.03\ \cancel{g}/mol} = 1.00\ mol$$

$$\begin{aligned}\text{Total number of moles} &= 0.500\ mol + 0.500\ mol + 1.50\ mol + 1.00\ mol \\ &= 3.50\ mol\end{aligned}$$

Applying Avogadro's law:

$$\frac{n_i}{V_i} = \frac{n_f}{V_f}\ \therefore\ V_f = \frac{3.50\ \cancel{mol} \times 22.4\ L}{1.00\ \cancel{mol}} = 78.4\ L$$

Therefore, the volume of the container is 78.4 L.

Check Your Solution

The molar volume at STP is 22.4 L/mol. Working with the numbers, $\frac{78.4\ L}{3.50\ mol} = 22.4$ L/mol. The results match.

10. Problem

(a) What volume do 2.50 mol of oxygen occupy at STP?
(b) How many molecules are present in this volume of oxygen?
(c) How many oxygen atoms are present in this volume of oxygen?

What Is Required?

(a) You need to find the volume of oxygen from the given number of moles.
(b) You need to find the number of molecules of this oxygen.
(c) You need to find the number of atoms present in this volume of oxygen.

What Is Given?

At STP, $V_i = 22.4$ L; $n_i = 1$ mol; $n_f = 2.50$ mol

You know Avogadro's number is 6.02×10^{23} molecules/mol. You know oxygen is a diatomic gas.

Plan Your Strategy

(a) Use Avogadro's law to solve for V_f.
(b) Multiply V_f by the Avogadro number to obtain the number of molecules of oxygen gas.

(c) Multiply the number of molecules of oxygen gas by 2 atoms to obtain the number of oxygen atoms in the sample.

Act on Your Strategy

(a) $\frac{n_i}{V_i} = \frac{n_f}{V_f} \therefore V_f = \frac{2.50 \text{ mol} \times 22.4 \cancel{\text{L}}}{1.00 \cancel{\text{L}}} = 56.0 \text{ L}$

(b) Number of molecules of oxygen $= 2.50 \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}}$

$= 1.5 \times 10^{24}$ molecules

(c) There are two oxygen atoms in one molecule of oxygen gas. Therefore, the number of oxygen atoms $= 2(1.5 \times 10^{24})$ molecules $= 3.01 \times 10^{24}$ atoms.

Check Your Solution

The molar volume is 22.4 L/mol. Working with the numbers, $\frac{56.0 \text{ L}}{2.50 \text{ mol}} = 22.4$ L/mol. The results match.

11. Problem

What volume do 2.00×10^{24} atoms of neon occupy at STP?

What Is Required?

You need to find the volume of the gas for the given number of moles.

What Is Given?

At STP, $V_i = 22.4$ L; $n_i = 1$ mol

Number of atoms $= 2.00 \times 10^{24}$

Avogadro's number is 6.02×10^{23} molecules/mol. Neon is a monatomic gas.

Plan Your Strategy

Neon is a monatomic so the number of atoms will equal the number of molecules of gas. Multiply the number of molecules by Avogadro's number to get the number of moles of gas as n_f. Use Avogadro's law and n_f to solve for V_f.

Act on Your Strategy

Number of moles of Ne $= \frac{2.00 \times 10^{24} \cancel{\text{molecules}}}{6.02 \times 10^{23} \cancel{\text{molecules}}/\text{mol}} = 3.32$ mol

$\frac{n_i}{V_i} = \frac{n_f}{V_f} \therefore V_f = \frac{3.32 \cancel{\text{mol}} \times 22.4 \text{ L}}{1.00 \cancel{\text{mol}}} = 74.4 \text{ L}$

Therefore, the volume of the 2.00×10^{24} atoms of neon at STP is 74.4 L.

Check Your Solution

The molar volume is 22.4 L/mol. Working with the numbers, $\frac{74.4 \text{ L}}{3.32 \text{ mol}} = 22.4$ L/mol. The results match.

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12. Problem

4.00 L of ammonia gas in a container holds 2.17 mol at 206 kPa. What is the temperature inside the container?

What Is Required?

You need to find the temperature of the container under the given parameters of the ammonia gas.

What Is Given?

$V = 4.00$ L

$P = 206$ kPa

$n = 2.17$ mol

R is assumed to be 8.314 kPa·L/mol·K, since P is given in kPa.

Plan Your Strategy
Rearrange the ideal gas law, $PV = nRT$, to solve for T.

Act on Your Strategy
$T = \frac{PV}{nR} = \frac{206\ \cancel{\text{kPa}} \times 4.00\ \cancel{\text{L}}}{2.17\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}\cdot\text{L}/\text{mol}}\cdot\text{K}} = 45.7\ \text{K}$
Therefore, the temperature in kelvin units is 45.7 K.

Check Your Solution
The units in the equation cancel to give the units of kelvins.
From the given values, the molar volume is $\frac{4\ \text{L}}{2.17\ \text{mol}} = 1.84\ \text{L/mol}$. This is much less than the 22.4 L/mol at STP. The result is expected since the temperature is much lower than standard temperature (273 K). This decreases the volume of the gas.

13. Problem
How many kilograms of chlorine gas are contained in 87.6 m^3 at 290 K and 2.40 atm?
Hint: 1 m^3 = 1000 L

What Is Required?
You need to find the mass of the chlorine gas under the given parameters.

What Is Given?
$V = 87.6\text{m}^3 = 87\ 600$ L (using the hint)
P = 2.40 atm
T = 290 K
R is assumed to be 0.08206 atm·L/mol·K, since P is given in atm.

Plan Your Strategy
Rearrange the ideal gas law, $PV = nRT$, to solve for n. Then multiply n by the molar mass of chlorine gas to obtain its mass.

Act on Your Strategy
$n = \frac{PV}{RT} = \frac{2.4\ \cancel{\text{atm}} \times 87600\ \cancel{\text{L}}}{0.08206\ \cancel{\text{atm}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 290\ \cancel{\text{K}}} = 8.834 \times 10^3\ \text{mol}$
Mass of chlorine gas $= n \times M$
$= (8.834 \times 10^3)\ \text{mol} \times 70.9\ \text{g/mol}$
$= 6.26 \times 10^5\ \text{g}$
Therefore, the mass of the chlorine gas in kilograms is 626 kg.

Check Your Solution
The units in the ideal gas equation cancel to give the units of moles.
From the given values, the molar volume is $\frac{87\ 600\ \text{L}}{8834\ \text{mol}} = 9.91\ \text{L/mol}$. This is almost half the 22.4 L/mol at STP. The result is expected since the pressure is more than double of standard pressure (1 atm). This would be expected to decrease the volume of the gas by almost half. Your answer is reasonable.

14. Problem
Calculate the volume of 3.03 g of hydrogen gas at a pressure of 560 torr and a temperature of 139 K.

What Is Required?
You need to find the volume of the hydrogen gas under the given parameters.

What Is Given?
P = 560 torr
T = 139 K
m = 3.03 g

R is assumed to be 62.37 mmHg·L/mol·K, since P is given in torr, which is numerically equivalent to mmHg.

Plan Your Strategy
First, calculate the number of moles by dividing the given mass by the molar mass of hydrogen gas. Then, rearrange the ideal gas law, $PV = nRT$, to solve for V. Remember to convert the pressure in torr to mmHg in order for the units to be able to cancel out.

Act on Your Strategy
P = 560 torr = 560 mmHg (since 760 mmHg = 760 torr)

$$n = \frac{m}{M} = \frac{3.03\ \cancel{g}}{2.02\ \cancel{g}/\text{mol}} = 1.5\ \text{mol}$$

$$V = \frac{nRT}{P} = \frac{1.5\ \cancel{\text{mol}} \times 62.37\ \cancel{\text{mmHg}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}} \times 139\ \cancel{\text{K}}}{560\ \cancel{\text{mmHg}}} = 23.2\ \text{L}$$

Check Your Solution
The units in the ideal gas equation cancel to give the units of moles.
$\frac{23.2\ \text{L}}{1.5\ \text{mol}}$ = 15.66 L/mol. This is somewhat lower than the molar volume (22.4 L/mol) at STP. The result is expected since the temperature in the question is much lower than standard temperature (273.15 K), so the volume would be decreased as well. Your answer is reasonable.

15. Problem
A 6.0×10^2 L reaction tank contains 5.0 mol of oxygen gas and 28 mol of nitrogen gas. If the temperature is 83°C, what is the pressure of the oxygen, in kPa?

What Is Required?
You need to find the pressure of the oxygen gas under the given parameters.

What Is Given?
$V = 6.0 \times 10^2$ L
T = 83°C
n = 5.0 mol oxygen gas
R is assumed to be 8.314 kPa·L/mol·K, since the answer for P must be in kPa.

Plan Your Strategy
Rearrange the ideal gas law, $PV = nRT$, to solve for P. Remember to convert the temperature in °C to degree kelvin in order for the units to be able to cancel out.

Act on Your Strategy
T = 83°C + 273.15 = 356 K

$$P = \frac{nRT}{V} = \frac{5.0\ \cancel{\text{mol}} \times 8.314\ \text{kPa}\cdot\cancel{\text{L}/\text{mol}\cdot\text{K}} \times 356\ \cancel{\text{K}}}{6.0 \times 10^2\ \cancel{\text{L}}} = 25\ \text{kPa}$$

Check Your Solution
The units in the ideal gas equation cancel to give the units of kPa.
$\frac{600\ \text{L}}{5\text{mol}}$ = 120 L/mol. This is much higher than the molar volume (22.4 L/mol) at STP. The result is expected since the pressure calculated is much less than standard pressure (101.325 kPa), while the temperature at 83°C is much higher than standard temperature (0°C). Both effects will result in a large increase in the volume of the gas. Your answer is reasonable.

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16. Problem

Oxygen makes up about 20% of our atmosphere. Find the density of pure oxygen gas, in g/L, for the conditions in the Sample Problem: 12.50°C and 126.63 kPa?

What Is Required?

You need to find the density of the pure oxygen gas at the give parameters.

What Is Given?

T = 12.50°C
P = 126.63 kPa
R is assumed to be 8.314 kPa·L/mol·K, since P is in kPa.
V is assumed to be 1.00 L.

Plan Your Strategy

Step 1 Convert the temperature in °C to K.
Step 2 Find the number of moles of oxygen by rearranging the ideal gas law to solve for *n*.
Step 3 Multiply *n* by the molar mass of oxygen to find its mass.
Step 4 Divide the mass by 1.0 L to obtain the density.

Act on Your Strategy

Step 1 $T = 12.50°\text{C} + 273.15 = 285.65 \text{ K}$
Step 2 $n = \frac{PV}{RT} = \frac{126.63 \cancel{\text{kPa}} \times 1.00 \cancel{\text{L}}}{8.314 \cancel{\text{kPa}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 285.65 \cancel{\text{K}}} = 0.0533 \text{ mol}$
Step 3 $m = n \times M = 0.0533 \cancel{\text{mol}} \times 32.00 \text{ g}/\cancel{\text{mol}} = 1.71 \text{ g}$
Step 4 The density of oxygen gas $= \frac{1.71 \text{ g}}{1.00 \text{ L}} = 1.71 \text{ g/L}$

Check Your Solution

When the units cancel out in the ideal gas equation, mol remains.
When the units cancel out in the density equation, g/L remain.

17. Problem

Find the density of butane gas, C_4H_{10} (in g/L) at SATP conditions: 298 K and 100 kPa.

What Is Required?

You need to find the density of the butane gas at the given parameters.

What Is Given?

T = 298 K
P = 100 kPa
R is assumed to be 8.314 kPa·L/mol·K, since P is in kPa.
V is assumed to be 1.00 L.

Plan Your Strategy

Step 1 Find the number of moles of butane by rearranging the ideal gas law to solve for *n*.
Step 2 Multiply *n* by the molar mass of butane to find its mass.
Step 3 Divide the mass by 1.0 L to obtain the density.

Act on Your Strategy

Step 1 $PV = nRT \therefore n = \frac{PV}{RT} = \frac{100 \cancel{\text{kPa}} \times 1.00 \cancel{\text{L}}}{8.314 \cancel{\text{kPa}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 298 \cancel{\text{K}}} = 0.0404 \text{ mol}$
Step 2 Mass of butane $= 0.0404 \cancel{\text{mol}} \times 58.14 \text{ g}/\cancel{\text{mol}} = 2.35 \text{ g}$
Step 3 Density of butane gas $= \frac{2.35 \text{ g}}{1.00 \text{ L}} = 2.35 \text{ g/L}$

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains.
When the units cancel out in the density equation, g/L remain.

18. **Problem**
The atmosphere of the imaginary planet Xylo is made up entirely of poisonous chlorine gas, Cl_2. The atmospheric pressure of this inhospitable planet is 155.0 kPa, and the temperature is 89°C. What is the density of the atmosphere?

What Is Required?
You need to find the density of the chlorine gas that makes up this planet's atmosphere.

What Is Given?
T = 89°C
P = 155.0 kPa
R is assumed to be 8.314 kPa·L/mol·K, since P is in kPa.
V is assumed to be 1.00 L.

Plan Your Strategy
Step 1 Convert the temperature in °C to K.
Step 2 Find the number of moles of chlorine gas by rearranging the ideal gas law to solve for n.
Step 3 Multiply n by the molar mass of Cl_2 to find its mass.
Step 4 Divide the mass by 1.0 L to obtain the density.

Act on Your Strategy
Step 1 $T = 89°\text{C} + 273 = 362\text{ K}$
Step 2 $n = \dfrac{PV}{RT} = \dfrac{155.0\ \cancel{\text{kPa}} \times 1.00\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa}{\cdot}\text{L}}/\text{mol}{\cdot}\cancel{\text{K}} \times 362\ \cancel{\text{K}}} = 0.0515\text{ mol}$
Step 3 Mass of chlorine gas $= n \times M = 0.0515\ \cancel{\text{mol}} \times 70.90\text{ g}/\cancel{\text{mol}} = 3.65\text{ g}$
Step 4 Density of the chlorine gas $= \dfrac{3.65\text{ g}}{1.00\text{ L}} = 3.65\text{ g/L}$

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains.
When the units cancel out in the density equation, g/L remain.

19. **Problem**
The atmosphere of planet Yaza, from the same star system as Xylo, is made of fluorine gas, F_2. The density of the atmosphere of Yaza is twice the density of the atmosphere on Xylo. The temperature of both planets is the same. What is the atmospheric pressure of Yaza?

What Is Required?
You need to find the atmospheric pressure of the planet Yaza.

What Is Given?
T = 362 K
D = 3.65 g/L × 2 = 7.30 g/L
The planet's gas is fluorine, F_2.
R is assumed to be 8.314 kPa·L/mol·K, to give P in kPa.

Plan Your Strategy
Step 1 Find the mass of fluorine in 1.00 L volume, using its density.
Step 2 Establish the number of moles of F_2 by dividing the mass by fluorine's molar mass.
Step 3 Use the ideal gas law to solve for P.

Act on Your Strategy

Step 1 Mass of fluorine gas in 1.00 L of atmosphere = 7.30 g/~~L~~ × 1.00 ~~L~~ = 7.30 g

Step 2 $n = \frac{m}{M} = \frac{7.30\ \cancel{g}}{38.00\ \cancel{g}/mol} = 0.192\ mol$

Step 3 $P = \frac{nRT}{V} = \frac{0.192\ \cancel{mol} \times 8.314\ kPa{\cdot}\cancel{L/mol{\cdot}K} \times 362\ \cancel{K}}{1.00\ \cancel{L}} = 578\ kPa$

Check Your Solution

When the units cancel out in the ideal gas equation, kPa remains.

Solutions for Practice Problems

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20. Problem

A 1.56 L gas sample has a mass of 3.22 g at 100 kPa and 281 K. What is the molar mass of the gas?

What Is Required?

You need to find the molar mass of the gas sample.

What Is Given?

V = 1.56 L
P = 100 kPa
m = 3.22 g
T = 281 K
R is assumed to be 8.314 kPa·L/mol·K, since P is in kPa.

Plan Your Strategy

Step 1 Find the number of moles of the gas sample using the ideal gas equation and solving for n.

Step 2 Divide the given mass by n to establish the molar mass of the gas.

Act on Your Strategy

Step 1 $n = \frac{PV}{RT} = \frac{100\ \cancel{kPa} \times 1.56\ \cancel{L}}{8.314\ \cancel{kPa{\cdot}L}/mol{\cdot}\cancel{K} \times 281\ \cancel{K}} = 0.0668\ mol$

Step 2 $M = \frac{m}{n} = \frac{3.22\ g}{0.0668\ mol} = 48.2\ g/mol$

Therefore, the molar mass of the gas sample is 48.2 g/mol.

Check Your Solution

When the units cancel out in the ideal gas equation, mol remains. The units of the molar mass are g/mol.

21. Problem

2.0 L of haloethane has a mass of 14.1 g at 344 K and 1.01 atm. What is the molar mass of haloethane?

What Is Required?

You need to find the molar mass of the haloethane.

What Is Given?

V = 2.0 L
P = 1.01 atm
m = 14.1 g
T = 344 K
R is assumed to be 0.08206 atm·L/mol·K, since P is in atm.

Plan Your Strategy

Step 1 Find the number of moles of the haloethane sample using the ideal gas equation and solving for n.

Step 2 Divide the given mass by n to establish the molar mass of the gas.

Act on Your Strategy

Step 1 $n = \frac{PV}{RT} = \frac{1.01\ \cancel{\text{atm}} \times 2.0\ \cancel{\text{L}}}{0.08206\ \cancel{\text{atm}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 344\ \cancel{\text{K}}} = 0.0716\ \text{mol}$

Step 2 $M = \frac{m}{n} = \frac{14.1\ \text{g}}{0.0716\ \text{mol}} = 2.0 \times 10^2\ \text{g/mol}$

Therefore, the molar mass of haloethane is 200 g/mol.

Check Your Solution

When the units cancel out in the ideal gas equation, mol remains. The units of the molar mass are g/mol.

22. Problem

A vapour has a mass of 0.548 g in 237 mL, at 373 K and 755 torr. What is the molar mass of the vapour?

What Is Required?

You need to find the molar mass of the vapour.

What Is Given?

V = 237 mL = 0.237 L
P = 755 torr = 755 mmHg
m = 0.548 g
T = 373 K
R is assumed to be 62.37 mmHg·L/mol·K, since P is in torr, which is numerically equivalent to mmHg.

Plan Your Strategy

Step 1 Find the number of moles of the vapour sample using the ideal gas equation and solving for n.
Step 2 Divide the given mass by n to establish the molar mass of the gas.

Act on Your Strategy

Step 1 $n = \frac{PV}{RT} = \frac{755\ \cancel{\text{mmHg}} \times 0.237\ \cancel{\text{L}}}{62.37\ \cancel{\text{mmHg}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 373\ \cancel{\text{K}}} = 0.00769\ \text{mol}$

Step 2 $M = \frac{m}{n} = \frac{0.548\ \text{g}}{0.00769\ \text{mol}} = 71.3\ \text{g/mol}$

Therefore, the molar mass of the vapour is 71.3 g/mol.

Check Your Solution

When the units cancel out in the ideal gas equation, mol remains. The units of the molar mass are g/mol.

23. Problem

The mass of a 5.00 L evacuated container is 125.00 g. When the container is filled with argon gas at 298 K and 105.0 kPa, it has a mass of 133.47 g.
(a) Calculate the density of argon under these conditions.
(b) What is the density of argon at STP?

What Is Required?

You need to find the density of argon at (a) the given conditions, and (b) under STP conditions.

What Is Given?

V = 5.00 L
P = 105.0 kPa
m of container alone = 125.00 g
m of container + argon = 133.47 g
T = 298 K

Standard T = 273 K
Standard P = 101.325 kPa

Plan Your Strategy

(a) The density is the mass of argon divided by the volume. First find the mass of argon by subtracting the final mass by the mass of the container.

(b) The volume of the gas at STP must be obtained. The combined gas law can be used to solve for V_{STP}.

Act on Your Strategy

(a) Mass argon = mass of filled container − mass of empty container
= 33.47 g − 125.00 g
= 8.47 g

Density $= \frac{m}{V} = \frac{8.47 \text{ g}}{5.00 \text{ L}} = 1.69$ g/L

Therefore, the density of argon gas at this temperature and pressure is 1.69 g/L.

(b) Volume occupied by argon at STP:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{105.0 \cancel{\text{kPa}} \times 5.00 \text{ L} \times 273 \cancel{\text{K}}}{298 \cancel{\text{K}} \times 101.3 \cancel{\text{kPa}}} = 4.748 \text{ L}$$

$$\text{Density} = \frac{8.47 \text{ g}}{4.748 \text{ L}} = 1.78 \text{ g/L}$$

Therefore, the density of argon gas at STP is 1.78 g/L.

Check Your Solution

For both (a) and (b), the units for density is g/L. When the units cancel out in the combined gas equation, L remains.

24. Problem

A gaseous compound contains 92.31% carbon and 7.69% hydrogen by mass. 4.35 g of the gas occupies 4.16 L at 22.0°C and 738 torr. Determine the molecular formula of the gas.

What Is Required?

You need to find the molecular formula of the gaseous compound.

What Is Given?

Carbon = 92.31% by mass
Hydrogen = 7.69% by mass
m = 4.35
V = 4.16 L
T = 22.0°C
P = 738 torr = 738 mmHg (since 1 torr = 1 mmHg)
Assume R = 62.37 mmHg·L/mol·K, since P is given in torr (equivalent to mmHg)

Plan Your Strategy

Step 1 Assume the total mass of the sample is 100.0 g to find the empirical formula. Multiply the mass percents of each element by 100 g to find the mass of each.

Step 2 Find the number of moles for each element by giving the mass by its molar mass. The mole ratio to its lowest whole number will give the empirical formula of the gas. Then, determine the molar mass of this empirical formula.

Step 3 Use the ideal gas equation to determine the number of moles of the gas.

Step 4 Divide the mass by the number of moles obtained in Step 3 to get the molar mass of the gas.

Step 5 Divide the calculated molar mass of the gas by the molar mass of the empirical formula. Multiply the ratio factor by the empirical formula to obtain the molecular formula.

Act on Your Strategy

Step 1 Mass of the carbon in the sample = 92.31% × 100.0 g = 92.31 g.
Mass of the hydrogen in the sample = 7.69% × 100.0 g = 7.69 g.

Step 2 $n(\text{carbon}) = \frac{m}{M} = \frac{92.31\ \cancel{\text{g}}}{12.01\ \cancel{\text{g}}/\text{mol}} = 7.69\ \text{mol}$

$n(\text{hydrogen}) = \frac{m}{M} = \frac{7.69\ \cancel{\text{g}}}{1.01\ \cancel{\text{g}}/\text{mol}} = 7.61\ \text{mol}$

The ratio of the elements in the compound is 1.0 mol of carbon to 1.0 mol of hydrogen. Therefore, the empirical formula of the unknown gas is CH. The molar mass of CH = 13.02 g/mol.

Step 3 $n = \frac{PV}{RT} = \frac{738\ \cancel{\text{mmHg}} \times 4.16\ \cancel{\text{L}}}{62.37\ \cancel{\text{mmHg}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 295\ \cancel{\text{K}}} = 0.169\ \text{mol}$

Step 4 $M = \frac{m}{n} = \frac{4.35\ \text{g}}{0.169\ \text{mol}} = 25.74\ \text{g/mol}$

Step 5 The ratio of the molar masses of the molecular formula:

empirical formula = $\frac{25.74\ \cancel{\text{g/mol}}}{13.02\ \cancel{\text{g/mol}}} = 2.00$

Therefore, the molecular formula of the unknown gas is twice the empirical formula. $CH \times 2 = C_2H_2$.

Check Your Solution

In step 3, the units to the ideal gas equation cancel out to give moles. C_2H_2 is an acceptable molecular formula; it is the hydrocarbon acetylene.

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25. Problem

Use the following balanced equation to answer the questions below.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

(a) What is the mole ratio of oxygen gas to water vapour?
(b) What is the volume ratio of oxygen gas to water vapour?
(c) What is the volume ratio of hydrogen gas to oxygen gas?
(d) What is the volume ratio of water vapour to hydrogen gas?

What Is Required?

(a) You need to find the mole ratio of oxygen to water vapour in the equation.
(b) You need to find the volume ratio of oxygen to water vapour.
(c) You need to find the volume ratio of hydrogen to oxygen.
(d) You need to find the volume ratio of water vapour to hydrogen.

What Is Given?

The balanced equation is given.

Plan Your Strategy

(a) The mole ratio can be determined from the whole number in front of each reactant and product in the balanced equation.
(b)–(d) The volume ratios are the same as the mole ratios, and can be assumed from the whole number in front of each reactant and product in the balanced equation.

Act on Your Strategy

(a) The mole ratio of oxygen gas to water vapour is 1:2
(b) The volume ratio of oxygen gas to water vapour is 1:2
(c) The volume ratio of hydrogen gas to oxygen gas is 2:1

(d) The volume ratio of water vapour to hydrogen gas is 2:2 or 1:1

Check Your Solution

First check that the balanced equation is correct. You can draw 2 diatomic hydrogen gas molecules and one diatomic oxygen molecule to confirm they fit together as 2 water molecules. The ratios are therefore correct.

26. Problem

1.5 L of propane gas are burned in a barbecue. The following equation shows the reaction. Assume all gases are at STP.

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

(a) What volume of carbon dioxide gas is produced?
(b) What volume of oxygen is consumed?

What Is Required?

(a) You need to find the volume of carbon dioxide produced when 1.5 L of propane gas is burned.
(b) You need to find the volume of oxygen consumed from the atmosphere when the propane is burned.

What Is Given?

The volume of propane burned is 1.5 L. You are given the balanced equation of the reaction.

Plan Your Strategy

(a) The volume ratios are the same as the mole ratios of the balanced equation. Use the ratio of C_3H_8 to CO_2 and the given volume of propane to equate and solve for the volume of carbon dioxide.
(b) Similarly, use the volume ratio of propane to oxygen, and the given volume of propane, to equate and solve for the volume of oxygen consumed.

Act on Your Strategy

(a) The mole ratio of C_3H_8:CO_2 is 1:3. Let x be the volume of carbon dioxide gas.

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} = \frac{x \text{ L } CO_2}{1.5 \text{ L } C_3H_8}$$

$$x = \frac{3 \cancel{\text{mol}}\ CO_2 \times 1.5 \text{ L } C_3H_8}{1 \cancel{\text{mol}}\ C_3H_8}$$

$$x = 4.5 \text{ L}$$

Therefore, the volume of CO_2 produced is 4.5 L.

(b) The mole ratio of C_3H_8:O_2 is 1:5. Let x be the volume of oxygen gas.

$$\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} = \frac{x \text{ L } O_2}{1.5 \text{ L } C_3H_8}$$

$$x = \frac{5 \cancel{\text{mol}}\ O_2 \times 1.5 \text{ L } C_3H_8}{1 \cancel{\text{mol}}\ C_3H_8}$$

$$x = 7.5 \text{ L}$$

Therefore, the volume of oxygen gas consumed from air is 7.5 L.

Check Your Solution

(a) The volume ratio of the C_3H_8:CO_2 is 1.5L : 4.5 L = 1 : 3 in the lowest ratio (dividing both by 1.5). This matches with the mole ratio of the balanced equation.
(b) The volume ratio of the C_3H_8:O_2 is 1.5L : 7.5 L = 1 : 5 in the lowest ratio (dividing both by 1.5). This matches with the mole ratio of the balanced equation.

27. Problem

Use the following equation to answer the questions below.

$$SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$$

(a) Balance the equation.

(b) 12.0 L of sulfur trioxide, $SO_{3(g)}$, are produced at 100°C. What volume of oxygen is consumed?

(c) What assumption must you make to answer part (b)?

What Is Required?

(a) You need to balance the given equation.

(b) You need to find the volume of oxygen gas consumed to produce 12.0 L of the sulfur trioxide.

(c) You need to list your assumptions used in calculating the answer to part (b).

What Is Given?

The reactants and products of the reaction are given. The volume of SO_3 produced is 12.0 L. The temperature is 100°C.

Plan Your Strategy

(a) Balance the equation by inspection.

(b) The mole ratio of the balanced equation is the same as the volume ratios. Use the ratio of O_2 to SO_3 and the given volume of SO_3 to equate and solve for the volume of oxygen.

Act on Your Strategy

(a) $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$

(b) $\frac{1 \text{ mol } O_2}{2 \text{ mol } SO_3} = \frac{x \text{ L } O_2}{12.0 \text{ L } SO_3}$

$x = \frac{1 \cancel{\text{mol}} O_2 \times 12.0 \text{ L } SO_3}{2 \cancel{\text{mol}} SO_3}$

$x = 6.00 \text{ L}$

Therefore, the volume of oxygen consumed is 6.00 L.

(c) You must assume that goes to completion.

Check Your Solution

(a) Use diagrams to confirm that 2 molecules of SO_2 and 1 molecule of diatomic oxygen gives rise to 2 molecules of SO_3.

(b) The volume ratio of the O_2:SO_3 is 6.00 L : 12.00 L = 1 : 2 in the lowest ratio (dividing both by 6.00). This matches with the mole ratio of the balanced equation.

28. Problem

2.0 L of gas A reacts with 1.0 L of gas B to produce 1.0 L of gas C. All gases are at STP.

(a) Write the balanced chemical equation for this reaction.

(b) Each molecule of gas A is made of two identical "a" atoms. That is, gas A is really $a_{2(g)}$. In the same way, each molecule of gas B is made of two identical "b" atoms. What is the chemical formula of gas C in terms of "a" and "b" atoms?

What Is Required?

(a) You need to give the balanced equation of this reaction.

(b) You need to find the chemical formula of the gas C.

What Is Given?

Gas A = 2.0 L; gas B = 1.0 L; gas C = 1.0 L. Both gases A and B are diatomic molecules.

Plan Your Strategy

(a) The volumes given can be considered as the volume ratios, which we know are the same as the mole ratios of a balanced equation.

(b) Using the mole ratio, and the fact that A and B are diatomic, you can "rebalance" the equation by manipulating the subscripts of the product, which must contain both the reactants.

Act on Your Strategy

(a) $2A + B \rightarrow C$

(b) $2a_{2(g)} + b_{2(g)} \rightarrow a_4b_{2(g)}$

Check Your Solution

Draw 2 diatomic "a" gas molecules and one diatomic "b" molecule to confirm they fit together as one a_4b_2 molecule. The ratios and formulas are therefore correct.

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29. Problem

Engineers design automobile air bags that deploy most instantly on impact. To do this, an air bag must provide a large amount of gas in a very short time. Many automobile manufacturers use solid sodium azide, NaN_3, along with suitable catalysts, to provide the gas that is needed to inflate the air bag. The balanced equation for this reaction is

$$2NaN_{3(g)} \rightarrow 2Na_{(s)} + 3N_{2(g)}$$

(a) What volume of nitrogen gas will be produced if 117.0 g of sodium azide are stored in the steering wheel at 20.2°C and 101.2 kPa?

(b) How many molecules of nitrogen are present in this volume?

(c) How many atoms are present in this volume?

What Is Required?

(a) You need to find the volume of N_2 gas produced from the stored amount of sodium azide.

(b) You need to find the number of N_2 molecules in this volume.

(c) You need to find the number of N atoms in this volume.

What Is Given?

m of sodium azide = 117.0 g

T = 20.2°C

P = 101.2 kPa

You are given the balanced equation for the reaction.

Assume R = 8.314 kPa·L/mol·K, since P is given in kPa.

You know Avogadro's number = 6.02×10^{23} molecules/mol.

Plan Your Strategy

(a) Follow the steps below:

Step 1 Divide the given mass of azide by its molar mass to determine the number of moles used.

Step 2 Use the mole ratio of NaN_3:N_2 from the balanced equation, and the number of moles NaN_3 calculated, to equate and solve for the number of moles of N_2.

Step 3 Use the ideal gas equation to solve for V by inputting the number of moles of N_2 calculated in step 2.

(b) Multiply the number of moles of nitrogen calculated in Step 2 by the Avogadro's constant to obtain the number of molecules.

(c) Nitrogen is a diatomic gas, meaning there are 2 N atoms for every molecule of nitrogen. Multiply the number of molecules calculated in part (b) by 2 to get the number of atoms in this volume.

Act on Your Strategy

(a) Step 1 Number of moles of (NaN_3) $= \dfrac{117.0 \text{ g}}{65.02 \text{ g/mol}} = 1.80$ mol

Step 2 Number of moles of $N_2 = \dfrac{3 \text{ mol } N_2}{2 \text{ mol } NaN_3} \times 1.80$ mol NaN_3 = 2.70 mol

Step 3 $V = \dfrac{nRT}{P} = \dfrac{2.70 \cancel{\text{mol}} \times 8.314 \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}} \times 293.2 \cancel{\text{K}}}{101.2 \cancel{\text{kPa}}} = 65.0 \text{ L}$

(b) 2.70 ~~mol~~ $N_2 \times (6.02 \times 10^{23})$ molecules/~~mol~~ $= 1.63 \times 10^{24}$ molecules

(c) The number of atoms is twice the number of molecules.
2 atoms/~~molecule~~ $\times (1.63 \times 10^{24})$ ~~molecule~~ $= 3.26 \times 10^{24}$ atoms

Check Your Solution

When the units cancel out in the ideal gas equation, L remains. The mole ratio of NaN_3:N_2 is 2:3, which is 1:1.5 in its lowest ratio. The ratio of the calculated mole amount, is $1.80 : 2.70 = 1 : 1.5$ in its lowest ratio (dividing both by 1.80). Your answer matches.

30. Problem

0.72 g of hydrogen gas, H_2, reacts with 8.0 L of chlorine gas, Cl_2, at STP. How many litres of hydrogen chloride gas, HCl, are produced?

What Is Required?

You need to find the volume of HCl produced in the reaction.

What Is Given?

m of H_2 = 0.72 g
V of Cl_2 = 8.0 L
T = 273 K
P = 101.3 kPa
Assume R = 8.314 kPa·L/mol·K, since P is assumed to be in kPa.

Plan Your Strategy

Step 1 Write down the balanced equation for the reaction.
Step 2 Divide the given mass of H_2 by its molar mass to determine the number of moles used.
Step 3 Use the mole ratio of H_2:HCl from the balanced equation, and the number of moles H_2 calculated, to equate and solve for the number of moles of HCl.
Step 4 Use the ideal gas equation to solve for V by inputting the number of moles of H_2 calculated in step 2.

Act on Your Strategy

Step 1 The balanced equation is: $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

Step 2 $n(H_2) = \dfrac{0.72 \cancel{\text{g}}}{2.02 \cancel{\text{g}}/\text{mol}} = 0.356$ mol

Step 3 n(HCl) = 2 × 0.356 mol = 0.713 mol

Step 4 $\therefore\ V = \dfrac{nRT}{P} = \dfrac{0.713 \cancel{\text{mol}} \times 8.314 \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}} \times 273 \cancel{\text{K}}}{101.3 \cancel{\text{kPa}}} = 16 \text{ L}$

Therefore, the volume of HCl produced is 16.0 L.

Check Your Solution

When the units cancel out in the ideal gas equation, L remains. The mole ratio of H_2:HCl is 1:2. The ratio of the calculated mole amounts, is 0.356:0.713, which is 1:2 in its lowest ratio (dividing both by 0.356). Your answer matches.

31. Problem

How many grams of baking soda (sodium hydrogen carbonate, $NaHCO_3$), must be used to produce 45 mL of carbon dioxide gas at 190°C and 101.3 kPa in a pan of muffins? (The mole ratio of $NaHCO_3$ to CO_2 is 1:1.)

What Is Required?
You need to find the mass of baking soda needed for the muffins.

What Is Given?
The mole ratio of $NaHCO_3$ to $CO_2 = 1:1$
V of $CO_2 = 45$ mL $= 0.045$ L
$T = 190°C$
$P = 101.3$ kPa
Assume $R = 8.314$ kPa·L/mol·K, since P is in kPa.

Plan Your Strategy
Step 1 Use the ideal gas equation to solve for the number of moles of CO_2 gas produced. Since the mole ratio of the two compounds is given as 1:1, then this calculated number of moles will be the same as that for the baking soda. (Remember to convert °C to K by adding 273.)
Step 2 Multiply the number of moles of baking soda, determined in step 1, by the molar mass of baking soda, to obtain its mass.

Act on Your Strategy
Step 1 $T = 190°C + 273 = 463$ K; $P = 101.3$ kPa, V of $CO_2 = 0.045$ L

$$n = \frac{PV}{RT} = \frac{101.3\ \cancel{\text{kPa}} \times 0.45\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa}}{\cdot}\cancel{\text{L}}/\text{mol}{\cdot}\cancel{\text{K}} \times 463\ \cancel{\text{K}}} = 0.00118 \text{ mol}$$

Therefore, the number of moles of $NaHCO_3$ is also 0.00118 mol.
Step 2 Mass of $NaHCO_3 = 0.00118\ \cancel{\text{mol}} \times 84.01\ \text{g}/\cancel{\text{mol}} = 9.9 \times 10^{-2}$ g

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains.
When the units cancel out in the mass determination, g remains.

32. Problem
How much zinc (in grams) must react with hydrochloric acid to produce 18 mL of gas at SATP? (**Hint**: Zinc chloride, $ZnCl_{2(s)}$ is a product.)

What Is Required?
You need to find the mass of zinc used in the reaction.

What Is Given?
V of the product is 18 mL $= 0.018$ L
Since conditions are SATP, then $P = 100$ kPa and $T = 298$ K.
Assume $R = 8.314$ kPa·L/mol·K, since P is in kPa.

Plan Your Strategy
Step 1 Write the balanced equation for the reaction to identify the gas produced.
Step 2 Use the ideal gas equation to solve for the number of moles of product gas produced.
Step 3 Use the mole ratio of Zn to product from the balanced equation, and the calculated number of moles of product from step 2, to equate and solve for the number of moles of Zn.
Step 4 Multiply the number of moles of Zn by its molar mass to obtain the mass in grams.

Act on Your Strategy
Step 1 The balanced equation is: $Zn_{(s)} + 2HCl_{(l)} \rightarrow ZnCl_{2(s)} + H_{2(g)}$
Step 2 $$n = \frac{PV}{RT} = \frac{100\ \cancel{\text{kPa}} \times 0.018\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa}}{\cdot}\cancel{\text{L}}/\text{mol}{\cdot}\cancel{\text{K}} \times 298\ \cancel{\text{K}}} = 0.000726 \text{ mol}$$
Step 3 The ratio of Zn to H_2 in the equation is 1:1, therefore the number of moles of Zn that must react $= 0.000726$ mol
Step 4 Mass of Zn $= 0.000726\ \cancel{\text{mol}} \times 65.39\ \text{g}/\cancel{\text{mol}} = 0.047$ g

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains.
When the units cancel out in the mass determination, g remains.

33. Problem
35 g of propane gas burned in a barbecue, according to the following equation:

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

All the gases are measured at SATP.
(a) What volume of water vapour is produced?
(b) What volume of oxygen is consumed?

What Is Required?
(a) You need to find the volume of $H_2O_{(g)}$ produced in the reaction.
(b) You need to find the volume of O_2 consumed in the reaction.

What Is Given?
m of propane = 35 g
Conditions are SATP, therefore P = 100 kPa and T = 298 K
The balanced equation is given.
Assume R = 8.314 kPa·L/mol·K, since P is in kPa.

Plan Your Strategy
(a) Follow the following steps.
Step 1 Divide the mass of propane by its molar mass to obtain its number of moles.
Step 2 Use the mole ratio of propane to water vapour, and the answer calculated in step 1, to equate and solve for the number of moles of $H_2O_{(g)}$.
Step 3 Apply the answer in Step 2 to the ideal gas equation to solve for the volume of water vapour produced.
(b) **Step 1** The mole ratio of propane gas to oxygen is 1:5, so multiply the number of moles of proane by 5 to get the number of moles of oxygen.
Step 2 Apply the answer in Step 2 to the ideal gas equation to solve for the volume of oxygen produced.

Act on Your Strategy
(a) **Step 1** Number of moles of propane $= n = \frac{m}{M} = \frac{35\text{ g}}{44.11\text{ g}} = 0.793$ mol
Step 2 Number of moles of water vapour = 4 × 0.793 mol = 3.172 mol
Step 3 $V = \frac{nRT}{P} = \frac{3.172\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}}\cdot\cancel{\text{K}} \times 298\ \cancel{\text{K}}}{100\ \cancel{\text{kPa}}} = 78.6\text{ L}$
Therefore, the volume of water vapour produced is 78.6 L.
(b) **Step 1** The number of moles of oxygen consumed = 5 × 0.793 mol = 3.965 mol.
Step 2 $V = \frac{nRT}{P} = \frac{3.965\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}}\cdot\cancel{\text{K}} \times 298\ \cancel{\text{K}}}{100\ \cancel{\text{kPa}}} = 98.2\text{ L}$
Therefore, the volume of oxygen consumed in the reaction is 98.2 L.

Check Your Solution
When the units cancel out in the ideal gas equations, L remains. The mole ratio of O_2:H_2O is 5:4, which is 1.25:1 in its lowest ratio. The ratio of the calculated mole amounts, is 3.965:3.172, which is 1.253:1 in its lowest ratio (dividing both by 3.172). Your answer is reasonable.

34. Problem
What mass of oxygen is reacted to produce 0.62 L of water vapour at 100°C and 101.3 kPa? Start by balancing the following equation:
$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$

What Is Required?
You need to find the mass of oxygen used in the reaction.

What Is Given?
V of $H_2O_{(g)}$ = 0.62 L
T = 100°C
P = 101.3 kPa
Assume R = 8.314 kPa·L/mol·K, since P is in kPa.
The reactants and products of the reaction are given.

Plan Your Strategy
Step 1 Balance the given equation.
Step 2 Use the ideal gas equation to solve for the number of moles of water vapour produced.
Step 3 Use the mole ratio of O_2 to $H_2O_{(g)}$ and the answer in Step 2 to equate and solve for the number of moles of oxygen used.
Step 4 Multiply the answer in Step 3 by the molar mass of oxygen gas to obtain its mass in grams.

Act on Your Strategy
Step 1 The balanced equation is: $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(g)}$
Step 2 Number of moles of water vapour

$$\therefore\ n = \frac{PV}{RT} = \frac{101.3\ \cancel{\text{kPa}} \times 0.62\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa}}{\cdot}\cancel{\text{L}}/\text{mol}{\cdot}\cancel{\text{K}} \times 373\ \cancel{\text{K}}} = 0.0203\ \text{mol}$$

Step 3 Number of moles of oxygen = $\frac{1}{2} \times 0.0203$ mol = 0.0101 mol
Step 4 Mass of oxygen = 0.0101 $\cancel{\text{mol}} \times 32.00\ \text{g}/\cancel{\text{mol}}$ = 0.32 g
Therefore, the mass of oxygen consumed in the reaction is 0.32 g.

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains. The mole ratio of O_2:H_2O is 0.5:1 or 1:2. The ratio of the calculated mole amounts, is 0.0101:0.0203, which is 1:2 in its lowest whole number ratio (dividing both by 0.0101). Your answer is reasonable.

Solutions for Practice Problems

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35. Problem
Oxygen, O_2, reacts with magnesium, Mg, to produce 243 g of magnesium oxide, MgO, at 101.3 kPa and 45°C. How many litres of oxygen are consumed? Start by writing the balanced equation.

What Is Required?
You need to find the volume of oxygen consumed in the reaction.

What Is Given?
The reactants and product of the reaction are given.
m of MgO = 243 g
T = 45°C
P = 101.3 kPa
Assume R = 8.314 kPa·L/mol·K, since P is in kPa.

Plan Your Strategy
Step 1 Write down the balanced equation.
Step 2 Divide the mass of MgO by its molar mass to obtain its number of moles.

Step 3 Use the mole ratio of O_2 to MgO in the balanced equation and the answer in Step 1 to equate and solve for the number of moles of O_2 used.
Step 4 Apply the ideal gas equation using *n* of O_2 to solve for the volume of oxygen consumed. Remember to convert °C to K by adding 273 to the give *T*.

Act on Your Strategy
Step 1 The balanced equation is: $O_{2(g)} + 2Mg_{(g)} \rightarrow 2MgO_{(s)}$
Step 2 Number of moles of MgO $= n = \frac{m}{M} = \frac{243\ \cancel{g}}{40.31\ \cancel{g}/mol} = 6.03$ mol
Step 3 Number of moles of O_2 produced = 6.03 mol / 2 = 3.01 mol
Step 4 $\therefore\ V = \frac{nRT}{P} = \frac{3.01\ \cancel{mol} \times 8.314\ \cancel{kPa}\cdot L/\cancel{mol\cdot K} \times 318\ \cancel{K}}{101.3\ \cancel{kPa}} = 79$ L
Therefore, the volume of oxygen gas reacted is 79 L.

Check Your Solution
When the units cancel out in the ideal gas equation, L remains. The mole ratio of O_2:MgO is 1:2. The ratio of the calculated mole amounts, is 3.01:6.03, which is 1:2 in its lowest whole number ratio (dividing both by 3.01). Your answer is reasonable.

36. Problem
Zinc reacts with nitric acid to produce 34 L of dry hydrogen gas at 900 torr and 20°C. How many grams of zinc are consumed?

What Is Required?
You need to find the mass of Zn used in the reaction.

What Is Given?
The reactants and products of the reaction are given.
V of H_2 = 34 L
P = 900 torr = 900 mmHg (since 1 torr = 1 mmHg)
T = 20°C
Assume R = 62.37 mmHg·L/mol·K, since P is in mmHg.

Plan Your Strategy
Step 1 Write down the balanced equation.
Step 2 Apply the ideal gas equation using V of H_2 to solve for the number of moles of H_2 produced. Remember to convert °C to K by adding 273 to the give *T*.
Step 3 Use the mole ratio of Zn to H_2 in the balanced equation and the answer in Step 2 to equate and solve for the number of moles of Zn used.
Step 4 Multiply the number of moles of Zn calculated in Step 3 to its molar mass to obtain its mass in grams.

Act on Your Strategy
Step 1 The balanced equation is: $2Zn_{(s)} + 2HNO_{3(l)} \rightarrow H_{2(g)} + 2ZnNO_{3(aq)}$
Step 2 Number of moles of hydrogen gas
$\therefore\ n = \frac{PV}{RT} = \frac{900\ \cancel{mmHg} \times 34\ \cancel{L}}{62.37\ \cancel{mmHg}\cdot\cancel{L}/mol\cdot\cancel{K} \times 293\ \cancel{K}} = 1.67$ mol
Step 3 Number of moles of zinc = 2 × 1.67 mol = 3.34 mol
Step 4 Mass of zinc consumed $= n \times M = 3.34\ \cancel{mol} \times 65.39\ g/\cancel{mol} = 218$ g

Check Your Solution
When the units cancel out in the ideal gas equation, mol remains. The mole ratio of Zn to H_2 is 2:1. The ratio of the calculated mole amounts, is 3.34:1.67, which is 2:1 in its lowest whole number ratio (dividing both by 1.67). Your answer is reasonable.

37. Problem
0.75 L of hydrogen gas is collected over water at 25.0°C and 101.6 kPa. What volume will the dry hydrogen occupy at 103.3 kPa and 25.0 C?

What Is Required?
The volume of the hydrogen at the new temperature and pressure is required.

What Is Given?
$P_{Total1} = 101.6$ kPa
$P_{Total2} = 103.3$ kPa
$T_1 = T_2 = 25.0°C$
$V_1 = 0.75$ L
$V_2 = ?$

Plan Your Strategy
Step 1 Determine the pressure of dry hydrogen at each total pressure given. P of water vapour is 3.17 kPa at 25°C, therefore P of hydrogen is P_{Total} minus P of water vapour.
Step 2 Use the combined gas law to solve for the V_2 of hydrogen. Remember to convert temperature to degree kelvins by adding 273 to the 25°C.

Act on Your Strategy
Step 1 P_1 exerted by hydrogen = 101.6 kPa − 3.17 kPa = 98.43 kPa
P_2 exerted by hydrogen = 103.3 kPa − 3.17 kPa = 100.13 kPa

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Step 2 $V_2 = \frac{P_1V_1T_2}{T_1P_2}$

$$V_2 = \frac{98.43\ \cancel{\text{kPa}} \times 0.75\ \text{L} \times 273\ \cancel{\text{K}}}{298\ \cancel{\text{K}} \times 101.13\ \cancel{\text{kPa}}}$$

$V_2 = 0.74$ L

Therefore, the volume of dry hydrogen at 103.3 kPa and 25°C is 0.74 L.

Check Your Solution
When the units cancel out in the combined gas equation, L remains. The pressure is slightly higher at 103.3 kPa, so you should expect the volume of gas to be slightly lower. This was indeed the case, so your result is reasonable.

38. Problem
3070 kg of coal burns to produce carbon dioxide. Assume that the coal is 95% pure carbon and the combustion is 80% efficient.
(**Hint**: The mole ration of $C_{(s)}$ to $CO_{2(g)}$ is 5:4.) How many litres of carbon dioxide are produced at SATP?

What Is Required?
You need to find the volume of CO_2 produced at SATP.

What Is Given?
m of coal = 3070 kg
Coal is 95% pure C
The reaction is 80% efficient.
Mole ratio of C to $CO_2 = 5:4$.
At SATP, $P = 100$ kPa and $T = 298$ K

Plan Your Strategy
Step 1 Determine the mass of C in coal by multiplying the mass of coal by the given mass percent.
Step 2 The reaction is only 80% efficient meaning only 80% of the total mass of C will go into CO_2. Multiply the mass of C by 80% to obtain its final mass used in the reaction.
Step 3 Divide the mass of C by its molar mass to obtain the number of moles.

Step 4 Use the mole ratio given and the number of moles of C to equate and solve for the number of moles of CO_2 produced.
Step 5 Use the ideal gas equation and input n of CO_2 to solve for the volume of CO_2.

Act on Your Strategy
Step 1 Mass of carbon in the coal sample = 3070 kg × 95%
= 2917 kg
= 2.917×10^6 g
Step 2 Actual mass of C used = (2.917×10^6) g × 80% = 2.334×10^6 g
Step 3 Number of moles of carbon = (2.334×10^6) g / 12.01 g/mol
= 1.94×10^5mol
Step 4 Number of moles of CO_2 produced is
4 mol / 5 mol × $(1.94 \times 10^5$ mol) = 1.55×10^5mol.
Step 5 $\therefore\ V = \frac{nRT}{P} = \frac{1.55 \times 10^5\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}} \times 298\ \cancel{\text{K}}}{100\ \cancel{\text{kPa}}} = 3.84 \times 10^6\ \text{L}$
Therefore, the volume of CO_2 produced in the 80% efficient reaction is 3.84×10^6 L.

Check Your Solution
When the units cancel out in the combined gas equation, L remains. Assume the reaction is 100% efficient:
Number of moles of C = $\frac{2.917 \times 10^6\ \cancel{\text{g}}}{12.01\ \cancel{\text{g}}/\text{mol}}$ = 2.43×10^5 mol
Number of moles of CO_2 = (4 mol / 5 mol) × (2.43×10^5) mol = 1.94×10^5 mol
Therefore, the mass of CO_2 produced = (1.94×10^5) mol × 44.01 g/mol.
= 8.538×10^6 g
= theoretical yield
Efficiency of the reaction was 80% = $\frac{\text{Actual yield}}{\text{Theoretical yield}}$
Actual yield is 8.538×10^6 g × 0.8 = 6.83×10^6 g.
Number of moles of CO_2 from step 4 = 1.55×10^5 mol.
Mass of this mol amount = $1.55 \times 10^5\ \cancel{\text{mol}} \times 44.01\ \text{g}/\cancel{\text{mol}} = 6.82 \times 10^6$ g. This is consistent with the actual yield calculated above.

39. Problem
When 7.48 g of iron reacts with chlorine gas, 21.73 g of product is formed.
(a) How many moles of chlorine are used?
(b) What is the formula for the product?
(c) Write the equation for the reaction that occurs.

What Is Required?
(a) You need to find the number of moles of chlorine used in the reaction.
(b) You need to find the formula of the product.
(c) You need to determine the balanced equation for the reaction that occurs.

What Is Given?
Fe = 7.48 g
Product = 21.73 g
Reactants are Fe and $Cl_{2(g)}$

Plan Your Strategy
We will assume the reaction proceeds to 100% efficiency. The law of conservation of mass determines that the mass of the product is equal to the mass of the reactants. The mass of chlorine is the mass of the product minus the mass of the Fe.
(a) The number of moles of Cl is the mass divided by its molar mass.

(b) **Step 1** Calculate the mass percents of each reactant in the product by dividing their mass by the mass of the product.

Step 2 Assume that 100 g of product is produced. The mass of each mass percent is then the mass percent expressed as a decimal multiplied by 100 g.

Step 3 Divide each mass by its own molar mass to get the mole amount.

Step 4 The mole ratio in its lowest whole number ratio gives the formula for the product.

(c) Knowing the formula of the product, you can now write down the equation and balance it.

Act on Your Strategy

(a) Mass of Cl = 21.73 g − 7.48 g = 14.25 g

Number of moles of $Cl_{2(g)} = \frac{14.25\ \cancel{g}}{70.90\ \cancel{g}/mol} = 0.201$ mol

(b) **Step 1** % Fe in product = $\frac{7.48\ g}{21.73\ g} \times 100\% = 34.4\%$

% Cl_2 in product = $\frac{14.25\ g}{21.73\ g} \times 100\% = 65.6\%$

Step 2 Assuming that 100 g of product is produced

Mass of Fe = 100 g × 0.344 = 34.4 g

Mass of Cl = 100 g × 0.656 = 65.6 g

Step 3 Number of moles Fe = $\frac{34.4\ \cancel{g}}{55.85\ \cancel{g}/mol} = 0.616$ mol

Number of moles Cl = $\frac{65.6\ \cancel{g}}{35.45\ \cancel{g}/mol} = 1.850$ mol

Step 4 The mole ratio shows a 1:3 ratio between Fe and Cl

Therefore, the formula must be $FeCl_3$.

(c) The balanced equation is: $2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$

Check Your Solution

Work backward with the balanced equation.

Number of moles of $FeCl_3$ produced = $\frac{21.75\ \cancel{g}}{162.2\ \cancel{g}/mol} = 0.134$ mol

The mole ratio of Fe to $FeCl_3$ is 2:2, which is 1:1. The number of moles of Fe used is also 0.134 mol. The mass of Fe used = 0.134 ~~mol~~ × 55.85 g/~~mol~~ = 7.48 g. This is the same amount as given in the question. Your answer matches.