**Name:** Key

# Chapter 13- Gas Packet

Gases and gas behavior is one of the most important and most fun things to learn during your year in chemistry. Over the years, I have developed, borrowed & stolen many different worksheets on gas behavior. So, in an effort to get rid of the question "Which worksheet are we working on?" I decided to put all of these gas worksheets in one packet. We will work through this packet and at the end; you will know all that you need to know regarding gas and gas behavior.

## *Section 1- Important terms*

1. What is the ***volume*** of a gas? What units are used to measure volume?

The amount of space a gas takes off. mL, L, $cm^3$

2. What is temperature? What units are used to measure temperature?

Temp is the average Kinetic energy of molecules. Measured in °C or Kelvin

Kelvin = °C + 273

3. What is ***pressure***? How is it calculated? What units are used to measure pressure?

Pressure is the force applied over an area. $P = \frac{Force}{Area}$ It is measured in atmospheres (atm), Kilopascals (KPa), mm Hg, torr and psi.

4. Using the following conversions- 1 atmosphere of pressure = 760 mm Hg = 760 torr = 101.3 kPa = 14.7 pounds per square inch (psi), convert the following pressure values.

a. 1 atm = <u>760</u> mm Hg

b. 0.550 atm = <u>55.7</u> kPa = <u>418</u> mm Hg

$$.550\,atm\left(\frac{101.3\,KPa}{1\,atm}\right) = 55.7\,KPa\left(\frac{760\,mmHg}{101.3\,KPa}\right) = 418\,mmHg$$

c. 18.375 psi = <u>1.25</u> atm = <u>126.6</u> kPa

$$18.375\,psi\left(\frac{1\,atm}{14.7\,psi}\right) = 1.25\,atm\left(\frac{101.3\,KPa}{1\,atm}\right) = 126.6\,KPa$$

## ***Section 2- The Kinetic Molecular Theory of Gases***

Earlier this year, we modeled the structure of matter and changes in matter using pictures, symbols or even using student actors. We discussed that molecules of matter move around each other—solids move the slowest, while gas molecules move very quickly. We also mentioned that gas molecules have a lot of space between them and that the molecules are not attracted to each other. These were all parts of the ***Kinetic Molecular Theory.***

Another way that matter can be modeled is through the use of ***analogies***. An ***analogy*** can help you compare certain features of an abstract idea or theory to a situation that is familiar to you. Read the analogy provided here and answer the questions that follow it concerning the kinetic theory of gases.

Imagine that a large group of dancers on an enclosed dance floor represents gas molecules bouncing around inside a container. The dancers move back and forth across the floor, but not off the floor.

1. Decide which of the four variables—volume, temperature, pressure or number of molecules—is most like each of the following choices. Explain your choices.
   a. the number of dancers

      ↳ # of molecules

   b. the size of the room

      ↳ volume

   c. the beat to the music

      ↳ temperature

   d. the number and force of collisions among the dancers

      ↳ pressure

2. How does each of the following situations relate to what you have learned about gases and the kinetic molecular theory?

   Temp molecules volume

   a. The beat of the music and the number of dancers remain the same, but the size of the dance floor increases.

      Temp ~~Pres~~ (-) Volume ↑
      ~~#~~ ~~do~~ # molecules (-) Pressure ↓ (less collisions)

   volume # of molecules Temp

   b. The size of the dance floor and the number of dancers remain the same, but the beat of the music becomes faster.

      Volume (-) Temp ↑
      # of molecules (-) Pressure ↑ (more collisions)

   volume Temp # of molecules

   c. The size of the dance floor and the beat of the music are kept the same, but the number of dancers increases.

      Volume (-) # of molecules ↑
      Temp (-) Pressure ↑ (more collisions)

   Temp # of molecules pressure

   d. The beat of the music gets faster, but the number of dancers and the number of collisions remain the same.

      # of molecules (-) Temp ↑
      Pressure (-) Volume ↑ (size of floor needs to get bigger)

***Section 3- The Gas Laws***

There are 3 mathematical relationships between the number of particles, the pressure, the volume and the temperature of a gas, also known as the gas laws. Outlined below are the 4 basic laws- Boyle's Law, Charles' Law & Gay-Lussac's Law.

I. **Boyle's Law**

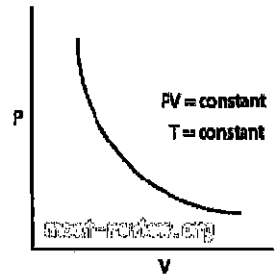
A. This law states that there is an inverse relationship between the amount of pressure applied to a gas and the volume that the gas occupies. **(P↑, V↓); (P↓, V↑)**

B. Temperature remains constant in this relationship.

C. The shape of a pressure-volume graph is the shape of a hyperbola.

D. The mathematical relationship for Boyle's Law is: $P_1V_1 = P_2V_2$

Boyle's Law




Example: A gas with a volume of 4.00 L of anesthetic gas changes from 105 kPa to 40.5 kPa. What will be the new volume of the gas?

$P_1 = 105\,kPa \quad V_1 = 4L$

$P_2 = 40.5\,kPa \quad V_2 = ?$

$P_1V_1 = P_2V_2$

$(105\,kPa)(4L) = (40.5\,kPa)(V_2)$

$\boxed{V_2 = 10.4\ L}$

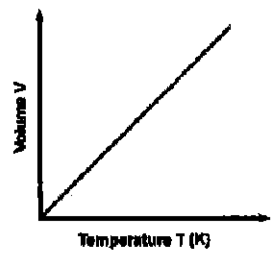
II. **Charles' Law**

A. This law states that there is a direct relationship between the temperature of a gas and the volume that the gas occupies. **(V↑, T↑); (V↓, T↓)**

B. Temperature must be converted into Kelvin (K = °C + 273)

C. Pressure exerted on the gas remains constant in this relationship.

D. The shape of a volume-temperature graph is a straight line.




E. The mathematical relationship for Charles' Law is: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Example: Exactly 5.00 L of air at -50 °C is warmed to 100 °C. What is the new volume of air if the pressure remains constant?

$V_1 = 5L$

$T_1 = -50°C + 273 = 223K$

$V_2 = ?$

$T_2 = 100°C + 273 = 373K$

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

$\frac{5L}{223K} = \frac{V_2}{373K}$

$\boxed{V_2 = 8.36L}$

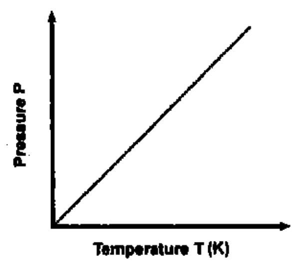
III. **Gay-Lussac's Law**

A. This law states that there is a direct relationship between the pressure that a gas exerts and the temperature of the gas. **(P↑, T↑); (P↓, T↓)**

B. Temperature must be converted into Kelvin (K = °C + 273)

C. The volume of the gas remains constant in this relationship.

D. The shape of a pressure-temperature graph is a straight line.




E. The mathematical relationship for Gay-Lussac's Law is: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Example: The pressure in an automobile tire is 198 kPa at 27 °C. At the end of a trip on a hot, sunny day, the pressure rose to 225 kPa. What is the temperature of the air in the tire in both Kelvin and degrees Celsius?

$P_1 = 198\,kPa$

$T_1 = 27°C + 273 = 300K$

$P_2 = 225\,kPa$

$T_2 = ?$

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

$\frac{198\,kPa}{300K} = \frac{225\,kPa}{T_2}$

$\boxed{T_2 = 340.9K = 67.9°C}$

## Basic Gas Law Problems

1. If $V_1 = 8260$ mL, $V_2 = 4.00$ L, and $P_1 = 724$ mm Hg, find the following:

   a. What is $V_1$ in L?

   $V_1 = 8.26L$

   b. What is $V_2$ in mL?

   $V_2 = 4000 mL$

   c. What is $P_1$ in atm?

   $P_1 = 724 mmHg\left(\frac{1atm}{760mmHg}\right) = \underline{.953atm}$

   d. What is $P_2$ in mm Hg?

   $P_1V_1 = P_2V_2$

   $(724mmHg)(8.26L) = P_2(4L)$

   $\boxed{P_2 = 1495 mmHg}$

   e. What is $P_2$ in atm?

   $P_1V_1 = P_2V_2$

   $(.953atm)(8.26L) = P_2(4L)$

   $P_2 = 1.97 atm$ (1.97 atm = 1495 mmHg)

2. A sample of gas occupies 2.50 L at 1.15 atm of pressure. What is the volume at standard atmospheric pressure? (standard pressure = 1 atm)

   $P_1 = 1.15atm$
   $V_1 = 2.50L$
   $P_2 = 1atm$
   $V_2 = ?$

   $P_1V_1 = P_2V_2$

   $(1.15atm)(2.50L) = (1atm)(V_2)$

   $\underline{V_2 = 2.875 L}$

3. A 500 mL sample of gas at 760 mm Hg is compressed to 100 mL. What is the new pressure?

   $P_1 = 760 mmHg$
   $V_1 = 500 mL$
   $P_2$ ?
   $V_2 = 100 mL$

   $P_1V_1 = P_2V_2$

   $(760 mmHg)(500mL) = P_2(100mL)$

   $\underline{P_2 = 3800 mL}$

4. A sample of gas occupies 45.2 mL at 720 mm Hg. What is its volume at standard atmospheric pressure (760 mm Hg)?

   $P_1 = 720 mmHg$
   $V_1 = 45.2 mL$
   $P_2 = 760 mmHg$
   $V_2 = ?$

   $P_1V_1 = P_2V_2$

   $(720 mmHg)(45.2mL) = (760mmHg)(V_2)$

   $\underline{V_2 = 42.8 mL}$

5. A sample of gas occupies 125 mL at standard pressure. What is its pressure when its volume is compressed to 75.0 mL?

   $P_1 = 1atm$
   $V_1 = 125 mL$
   $P_2 = ?$
   $V_2 = 75 mL$

   $P_1V_1 = P_2V_2$

   $(1atm)(125mL) = P_2(75mL)$

   $\underline{P_2 = 1.67 atm}$

6. A steel cylinder contains 2 cubic feet of a gas under a pressure of 1000 psi at room temperature, 20 °C. More heat is added until the pressure inside the cylinder is 1500 psi. What is the new temperature in the cylinder? Give your answer in both Kelvin and Celsius. Assume the volume remains constant.

   $P_1 = 1000 psi$
   $T_1 = 20°C = 293K$
   $P_2 = 1500 psi$
   $T_2 = ?$

   $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

   $\frac{1000psi}{293K} = \frac{1500psi}{T_2}$

   $\boxed{T_2 = 439.5K = 166.5°C}$

7. The pressure of a sample of argon gas is 657 mm Hg when the temperature is 10 °C. If the Kelvin temperature is doubled, what is the new pressure?

$P_1 = 657\ mmHg$ $\quad P_2 = ?$

$T_1 = 10°C = 283K$ $\quad T_2 = 2(283) = 566K$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{657\ mmHg}{283K} = \frac{P_2}{566K}$$

$$\boxed{P_2 = 1314\ mmHg}$$

8. A given mass of hydrogen chloride gas occupies a volume of 50.0 mL at 20 °C. Determine its volume at 127 °C, pressure remaining constant.

$V_1 = 50\ mL$ $\quad V_2 = ?$

$T_1 = 20°C = 293K$ $\quad T_2 = 127°C = 400K$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{50mL}{293K} = \frac{V_2}{400K}$$

$$\boxed{V_2 = 68.3\ mL}$$

9. A mass of hydrogen gas occupies 100 mL at a temperature of -73 °C. At what temperature will this same mass occupy 150 mL?

$V_1 = 100\ mL$ $\quad V_2 = 150\ mL$

$T_1 = -73°C = 200K$ $\quad T_2 = ?$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{100\ mL}{200K} = \frac{150\ mL}{T_2}$$

$$\boxed{T_2 = 300K}$$

10. A gas is kept in a steel cylinder under a pressure of 35 lb/in$^2$. Its temperature is 27 °C. The gas is then allowed to cool to room temperature of 20 °C. What is the new pressure?

$P_1 = 35\ psi$ $\quad P_2 = ?$

$T_1 = 27°C = 300K$ $\quad T_2 = 20°C = 293K$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{35\ psi}{300K} = \frac{P_2}{293K} \qquad \boxed{P_2 = 34.2\ psi}$$

<u>The Combined Gas Law</u>- Used when 2 variables are changing like pressure & temperature.

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

11. A gas has a volume of 150 mL at a pressure of 700 mm Hg and a temperature of 27 °C. What volume will the gas occupy at 0 °C and 1 atmosphere?

$V_1 = 150\ mL$ $\quad V_2 = ?$

$P_1 = 700\ mmHg$ $\quad P_2 = 1atm = 760\ mmHg$

$T_1 = 27°C = 300K$ $\quad T_2 = 0°C = 273K$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(700\ mmHg)(150mL)}{300K} = \frac{(760\ mmHg)(V_2)}{273K}$$

$$\boxed{V_2 = 125.7\ mL}$$

12. If 300 mL of a gas at 6 atm or pressure is heated from 227 °C to 333 °C and the pressure decreased to 3 atm, what is the new volume?

$V_1 = 300\ mL$ $\quad V_2 = ?$

$P_1 = 6atm$ $\quad P_2 = 3atm$

$T_2 = 227°C = 500K$ $\quad T_2 = 333°C = 606K$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(6atm)(300mL)}{500K} = \frac{(3atm)(V_2)}{606K}$$

$$\boxed{V_2 = 727.2\ mL}$$

13. When the temperature is 17 °C and the pressure is 650 mm Hg, a gas has a volume of 220 mL. Conditions are now changed so that the volume is now 400 mL and the pressure is 680 mm Hg. What is the new temperature?

$P_1 = 650\,\text{mmHg}$ $\quad P_2 = 680\,\text{mmHg}$

$V_1 = 220\,\text{mL}$ $\quad V_2 = 400\,\text{mL}$

$T_1 = 17°C = 290K$ $\quad T_2 = ?$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(650\,\text{mmHg})(220\,\text{mL})}{290K} = \frac{(680\,\text{mmHg})(400\,\text{mL})}{T_2}$$

$$\boxed{T_2 = 551.6K = 278.6°C}$$

14. The volume of a gas is 500 mL at 2 atm pressure and 0 °C. The volume becomes 250 mL when the temperature is raised to 546 °C. What is the new pressure?

$P_1 = 2\,\text{atm}$ $\quad P_2 = ?$

$T_1 = 0°C = 273K$ $\quad T_2 = 546°C = 819K$

$V_1 = 500\,\text{mL}$ $\quad V_2 = 250\,\text{mL}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(2\,\text{atm})(500\,\text{mL})}{273K} = \frac{P_2(250\,\text{mL})}{819K}$$

$$\boxed{P_2 = 12\,\text{atm}}$$

15. Find the volume of a gas at standard conditions (1 atm and 0 °C) when a gas occupies 100 mL when the temperature is -23°C and a pressure of 700 mm Hg.

$V_1 = 100\,\text{mL}$ $\quad V_2 = ?$

$T_1 = -23°C = 250K$ $\quad T_2 = 273K$

$P_1 = 700\,\text{mmHg}$ $\quad P_2 = 760\,\text{mmHg}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(700\,\text{mmHg})(100\,\text{mL})}{250K} = \frac{(760\,\text{mmHg})(V_2)}{273K}$$

$$\boxed{V_2 = 100.6\,\text{mL}}$$

<u>The Ideal Gas Law</u>

Pressure — Number of moles — Temperature

$$PV = nRT$$

Volume — Gas constant

$$R = 0.08206\frac{L \cdot atm}{mol \cdot K} = .0821\frac{L \cdot atm}{mol \cdot K}$$

Differences between Real & Ideal Gases

Ideal Gases have no attraction between the particles, where the real gases have a slight attraction.

16. How many moles of gas does it take to occupy 120 L at a pressure of 2.3 atmospheres and a temperature of 67 °C?

$n = ?$

$V = 120\,L$ $\quad T = 67°C = 340\,K$

$P = 2.3\,\text{atm}$ $\quad R = .0821\,\frac{L \cdot atm}{mol \cdot K}$

$$PV = nRT$$

$$(2.3\,\text{atm})(120\,L) = n\left(.0821\,\frac{L\cdot atm}{mol\cdot K}\right)(340K)$$

$$\boxed{n = 9.89 \text{ moles gas}}$$

17. If I have a 50 L container that holds 45 moles of gas at a temperature of 200 °C, what is the pressure inside the container?

$P = ?$

$V = 50\,L$

$n = 45$ moles

$T = 200°C = 473\,K$

$R = .0821\,\frac{L\cdot atm}{mol \cdot K}$

$$PV = nRT$$

$$P(50\cancel{L}) = (45\,\cancel{\text{moles}})\left(.0821\,\frac{\cancel{L}\cdot atm}{\cancel{mol\cdot K}}\right)(473\cancel{K})$$

$$\boxed{P = 34.9\,\text{atm}}$$

18. It is not safe to put aerosol canisters in a campfire, because the pressure inside the canister gets very high and they can explode. If I have a 1.0 L canister that holds 2 moles of gas, and the campfire temperature is 1400 °C, what is the pressure inside the canister?

P=?
V = 1 L
n = 2 moles
T = 1400°C = 1673 K
R = .0821 $\frac{L \cdot atm}{mol K}$

$PV = nRT$

$P(1L) = (.0821 \frac{L \cdot atm}{mol K})(2 moles)(1673K)$

$\boxed{P = 274.7 \text{ atm}}$

19. SCUBA divers use tanks with pressurized air to help them breathe underwater. If 243.6 moles of air are compressed in a tank to a pressure of 200 atms and a temperature of 27 °C, what size container is the diver carrying?

R = .0821 $\frac{L \cdot atm}{mol K}$
n = 243.6 moles
P = 200 atm
T = 27°C = 300 K
V = ?

$PV = nRT$

$(200 atm)(V) = (243.6 moles)(.0821 \frac{L \cdot atm}{mol K})(300K)$

$\boxed{V = 30.0 L}$

20. I have a balloon that can hold 100 L of air. If I blow up this balloon with 3 moles of oxygen gas at a pressure of 1 atmosphere, what is the temperature of the balloon?

V = 100 L
n = 3 moles
P = 1 atm
T = ?
R = .0821 $\frac{L \cdot atm}{mol K}$

$PV = nRT$

$(1 atm)(100 L) = (3 moles)(.0821 \frac{L \cdot atm}{mol K})(T)$

$\boxed{T = 406 K = 133°C}$

21. What volume does 1 mole of a gas occupy when it is under standard conditions?

V = ?
n = 1 mole
P = 1 atm
T = 273 K
R = .0821 $\frac{L \cdot atm}{mol K}$

$PV = nRT$

$(1 atm)(V) = (1 mole)(.0821 \frac{L \cdot atm}{mol K})(273K)$

$\boxed{V = 22.4 L}$

1 mole of any gas at standard conditions occupies 22.4 L

22. A physical chemist measured the lowest pressure achieved in the laboratory – about 1.32 x $10^{-18}$ atm. How many moles of gas are present in a 1.00 L sample at that pressure if the sample's pressure is 22 °C?

P = $1.32 \times 10^{-18}$ atm
R = .0821
n = ?
V = 1 L
T = 22°C = 295 K

$PV = nRT$

$(1.32 \times 10^{-18} atm)(1L) = n(.0821 \frac{L \cdot atm}{mol K})(295K)$

$\boxed{n = 5.45 \times 10^{-20} \text{ moles}}$

23. How many moles are contained in 13.5 liters of gas in a helium balloon under standard conditions?

P = 1 atm
V = 13.5 L
R = .0821 $\frac{L \cdot atm}{mol K}$
n = ?
T = 273 K

$PV = nRT$

$(1 atm)(13.5 L) = n(.0821 \frac{L \cdot atm}{mol K})(273K)$

$\boxed{n = .602 \text{ moles}}$

or

$13.5 \cancel{L} \left(\frac{1 mol}{22.4 \cancel{L}}\right) = \boxed{.602 \text{ moles}}$

***Using the ideal gas law***

The ideal gas law is PV = nRT. We can use the ideal gas law to determine the density of a gas and even the molar mass of a gas. We know that moles, n, can be calculated using the formula: n = mass/ molar mass. We also know that density, D, can be calculated using the formula: d = mass/volume. Using these 2 formulas and substituting them into the PV = nRT formula, you can have the following 2 formulas (which can be found on page 456 of your book:

**Molar Mass, MM =** $\frac{(\text{mass} \times R \times T)}{(V \times P)} = \frac{(D \times R \times T)}{P}$ **Density =** $\frac{(P \times MM)}{(R \times T)}$

Using the ideal gas law in addition to these formulas, complete the following problems.

24. Propane ($C_3H_8$) is a gas commonly used as a home fuel for cooking and heating. Calculate the volume that 0.540 moles of propane occupies at STP. Think about the size of this volume and the amount of propane that it contains. Why do you think propane is usually liquefied before it is transported?

$n = .540$ moles
$P = 1$ atm
$T = 273$ K
$R = .0821 \frac{L \cdot atm}{mol \cdot K}$
$V = ?$

$PV = nRT$

$(1 atm)(V) = (.540 moles)(.0821 \frac{L \cdot atm}{mol \cdot K})(273 K)$

$\boxed{V = 12.1 \text{ L}}$

It is usually liquefied so the propane will occupy a smaller volume, making it easier to ship.

25. 10.72 g of propane gas is in a 3 L tank under 2 atm of pressure and a temperature of 300 K. What is the molar mass of this gas?

mass = 10.72 g
$V = 3 L$
$P = 2 atm$
$T = 300 K$
$R = .0821 \frac{L \cdot atm}{mol \cdot K}$
$MM = ?$

$MM = \frac{\text{mass} \times R \times T}{V \times P} = \frac{(10.72 g \times .0821 \frac{\cancel{L \cdot atm}}{mol \cdot \cancel{K}} \times 300 \cancel{K})}{(3 \cancel{L} \times 2 \cancel{atm})}$

$\boxed{MM = 44.0 \text{ g/mole}}$

$C_3H_8$
$(3 \times 12) + (8 \times 1) = 44$

26. Geraniol is a compound found in rose oil that is used in perfumes. What is the molar mass of geraniol if its vapor has a density of 0.480 g/L at a temperature of 260 °C and a pressure of 0.140 atm?

$D = .480$ g/L
$T = 260°C = 533 K$
$P = .140$ atm
$R = .0821 \frac{L \cdot atm}{mol \cdot K}$
$MM = ?$

$MM = \frac{D \times R \times T}{P} = \frac{(.480 g/\cancel{L} \times .0821 \frac{\cancel{L \cdot atm}}{mol \cdot \cancel{K}} \times 533 \cancel{K})}{.140 \cancel{atm}}$

$\boxed{MM = 150.0 \text{ g/mole}}$

27. Determine the density of chlorine gas ($Cl_2$) at 22.0 °C and 1.00 atm.

$Cl_2$

$2 \times 35.5 = 71\ g/mole$

P = 1 atm

T = 22°C = 295K

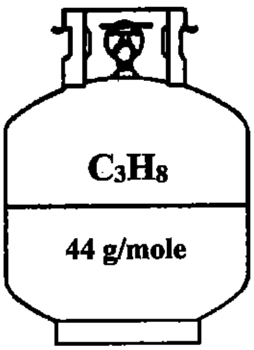
R = .0821 $\frac{L\ atm}{mol\ K}$

MM = 71 g/mole

D = ?

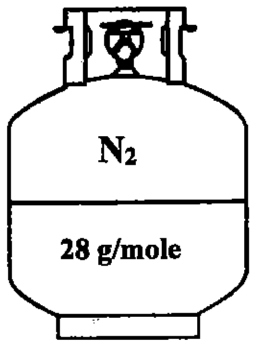
$$D = \frac{(P \times MM)}{(R \times T)} = \frac{((1\ atm) \times (71\ g/mole))}{(.0821 \frac{L\ atm}{mol\ K} \times 295K)} = \boxed{2.93\ g/L}$$

Two answer questions # 27 & 28, consider two tanks of gas. Tank 1 has 520 grams of $C_3H_8$ and Tank 2 has 380 grams of $N_2$.




520g

$n = \frac{520}{44} = 11.82\ mole$




380g

$n = \frac{380}{28} = 13.57\ moles$

28. Which of the gases described above occupies the greatest volume at STP? Explain.

→ $N_2$ has a greater volume. More moles

$C_3H_8$

P = 1 atm

T = 273K

R = .0821 $\frac{L\ atm}{mol\ K}$

V = ?

n = 11.82 moles

PV = nRT

(1 atm)(V) = (11.82 moles)(.0821)(273K)

$\boxed{V = 264.9\ L}$

– or –

$11.82\ moles \times 22.4 \frac{L}{mole} = \boxed{264.7L}$

$N_2$

P = 1 atm

V = ?

n = 13.57 moles

R = .0821 $\frac{L\ atm}{mol\ K}$

T = 273K

PV = nRT

(1 atm)(V) = (13.57)(.0821)(273)

$\boxed{V = 304.2\ L}$

–or–

$13.57\ moles \times 22.4 \frac{L}{mole} = \boxed{304L}$

29. If the tanks described above each hold 4.00 L, what is the pressure inside each tank at standard temp (273 K)?

$C_3H_8$

P = ?

V = 4L

n = 11.82 moles

R = .0821 $\frac{L\ atm}{mol\ K}$

T = 273K

PV = nRT

P(4) = (11.82)(.0821)(273)

$\boxed{P = 66.2\ atm}$

$N_2$

P = ?

V = 4L

n = 13.57 moles

R = .0821 $\frac{L\ atm}{mol\ K}$

T = 273K

PV = nRT

P(4) = (13.57)(.0821)(273)

$\boxed{P = 76.0\ atm}$

***Section 4- Other Laws involving gases***

Dalton's Law of Partial Pressures

This law states that if there is a mixture of gases (like the atmosphere); the total pressure the gas mixture exerts is equal to the sum of the pressures of each individual gas. There are 2 formulas:

**$P_T = P_A + P_B + P_c + \ldots$** **$P_T$ = total pressure; $P_A$ = pressure of gas A (and so on)**

**$P_A = (\% A)P_T$** **% A = percentage of gas A in mixture**

30. A metal tank contains 3 gases, oxygen, helium & nitrogen. If the partial pressures of the three gases in the tank are 35 atm of $O_2$, 5 atm of $N_2$, and 25 atm of He, what is the total pressure in the tank?

$$P_T = P_{O_2} + P_{N_2} + P_{He} = 35\,\text{atm} + 5\,\text{atm} + 25\,\text{atm} = \boxed{65\,\text{atm}}$$

31. Blast furnaces give off many unpleasant and unhealthy gases. If the total air pressure is 752.4 mm Hg, the partial pressure of carbon dioxide is 5.065 kPa, and the partial pressure of hydrogen sulfide is 0.02 atm, what is the partial pressure of the remaining air?

$$5.065\,\text{kPa}\left(\frac{760\,\text{mmHg}}{101.3\,\text{kPa}}\right) = 38\,\text{mmHg}$$

$$.02\,\text{atm}\left(\frac{760\,\text{mmHg}}{1\,\text{atm}}\right) = 15.2\,\text{mmHg}$$

$$P_T = P_{CO_2} + P_{HS} + P_{air}$$

$$752.4\,\text{mmHg} = 38\,\text{mmHg} + 15.2\,\text{mmHg} + P_{air}$$

$$\boxed{P_{air} = 699.2\,\text{mmHg}}$$

- all units need to be the same.

32. If the air from problem #2 contains 22% oxygen, what is the partial pressure of oxygen in air near a blast furnace?

$$P_{O_2} = P_{air} \times .22 = (699.2\,\text{mmHg}) \times .22 = \boxed{153.8\,\text{mmHg}}$$

33. Air is a mixture of gases. By percentage, it is roughly 78 % nitrogen, 21 % oxygen and 1 % argon. (There are other trace amounts of many other gases in air.) If the atmospheric pressure is 760 mm Hg, what are the partial pressures of nitrogen, argon and oxygen in the atmosphere?

$$P_{N_2} = P_T \times .78 = 760\,\text{mmHg} \times .78 = \boxed{592.8\,\text{mmHg}}$$

$$P_{O_2} = P_T \times .21 = 760\,\text{mmHg} \times .21 = \boxed{159.6\,\text{mmHg}}$$

$$P_{Ar} = P_T \times .01 = 760\,\text{mmHg} \times .01 = \boxed{7.6\,\text{mmHg}}$$

Add up to 760

Graham's Law

Gases spread out when they move around. This is evident when you smell a skunk. It stinks at first, but then the smell dissipates. The lighter gases (or less dense gases) move quicker than the heavier gases. This is known as Graham's Law. The formula is:

$$\frac{\text{Rate of effusion of gas}_1}{\text{Rate of effusion of gas}_2} = \sqrt{\frac{M_2}{M_1}} \quad \text{M = molar mass of each gas}$$

34. Under the same conditions of temperature and pressure, how many times faster will hydrogen ($H_2$) effuse (move) compared to carbon dioxide ($CO_2$)?

$$\frac{\text{rate } H_2}{\text{rate } CO_2} = \sqrt{\frac{MM_{CO_2}}{MM_{H_2}}} = \sqrt{\frac{44\text{ g/mol}}{2\text{ g/mol}}} = 4.69$$

$\boxed{H_2 \text{ is } 4.69\times \text{ faster than } CO_2}$

35. If the carbon dioxide in the previous problem takes 32 seconds to effuse, how long will the hydrogen take?

$$\frac{t_{CO_2}}{t_{H_2}} = \frac{\text{rate } H_2}{\text{rate } CO_2} = 4.69 \qquad t_{CO_2} = 32\text{ sec} \qquad \frac{32\text{ sec}}{t_{H_2}} = 4.69$$

$\boxed{t_{H_2} = 6.82\text{ sec}}$

36. What is the relative rate of diffusion of $NH_3$ compared to He? Does $NH_3$ move faster or slower than He? Slower

$$\frac{\text{rate } He}{\text{rate } NH_3} = \sqrt{\frac{MM_{NH_3}}{MM_{He}}} = \sqrt{\frac{17\text{ g/mol}}{4\text{ g/mol}}} = 2.06$$

$\boxed{He \text{ is } 2.06\times \text{ faster than } NH_3}$

37. If the He in the previous problem takes 20 seconds to effuse, how long will the $NH_3$ take?

$$\frac{t_{NH_3}}{t_{He}} = \frac{\text{rate } He}{\text{rate } NH_3} \qquad \frac{t_{NH_3}}{20\text{ sec}} = 2.06$$

$\boxed{t_{NH_3} = 41.2\text{ sec}}$

38. An unknown gas diffuses 0.25 times as fast as He. What is the molecular mass of the unknown gas?

$$\frac{\text{rate}_{Gas}}{\text{rate } He} = .25 \qquad \frac{\text{rate}_{Gas}}{\text{rate } He} = \sqrt{\frac{MM_{He}}{MM_{Gas}}} \qquad .25 = \sqrt{\frac{4}{MM_{Gas}}}$$

$$.0625 = \frac{4}{MM_{Gas}} \rightarrow \boxed{MM_{Gas} = 64\text{ g/mol}}$$

39. Two porous containers are filled with hydrogen and neon respectively. Under identical conditions, 2/3 of the hydrogen escapes in 6 hours. How long will it take for half of the neon to escape?

$H_2$ 2/3 of $H_2$ escapes in 6 hours, so all of $H_2$ escapes in 9 hours

Ne

$$\frac{t_{Ne}}{t_{H_2}} = \sqrt{\frac{MM_{Ne}}{MM_{H_2}}} \qquad \frac{t_{Ne}}{9\text{ hr}} = \sqrt{\frac{20\text{ g}}{2\text{ g}}}$$

$t_{Ne} = 28.5$ hours for all of Neon

it takes 14.23 hours for half of Ne to escape.

***Section 5- Gas Stoichiometry***

Just like stoichiometry involving solids, you can compare gaseous reactants and products in mole ratios. You would use the ideal gas law to calculate the # of moles of a gas. You would then use your mole ratios to get to what the question is asking.

However, if you are comparing 2 gases that are existing under identical pressure and temperature conditions, you can use the mole ratio in the balanced equation to compare volumes of reactant and products.

40. How many liters of propane gas ($C_3H_8$) will undergo complete combustion with 34.0 L of oxygen gas?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

$$34\,L\,O_2\left(\frac{1\ C_3H_8}{5\ O_2}\right) = \boxed{6.8\,L\,C_3H_8}$$

41. How many liters of oxygen ($O_2$) is needed to completely combust 2.36 L of methane gas ($CH_4$)?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

$$2.36\,L\,CH_4\left(\frac{2\ O_2}{1\ CH_4}\right) = \boxed{4.72\,L\,O_2}$$

42. Ammonium nitrate is a common ingredient in chemical fertelizers. Use the reaction shown to calculate the mass of solid ammonium nitrate that must be used to obtain 0.100 L of dinitrogen oxide gas at STP?

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

P = 1 atm
V = .1 L
n = ?
R = .0821 $\frac{L\cdot atm}{mol\cdot K}$
T = 273 K

$PV = nRT$

$(1)(.1) = n(.0821)(273)$

$n = .00446\ moles\ N_2O \left(\frac{1\ mol\ NH_4NO_3}{1\ mol\ N_2O}\right)\left(\frac{80\,g}{1\ mol\ NH_4NO_3}\right) = \boxed{.357\,g\ NH_4NO_3}$

43. Refer to the balanced equation in question #40 to calculate the mass of water vapor is formed when 3.00 L of propane gas is completely combusted to form water vapor and carbon dioxide at 350 °C and 0.990 atm.

Propane
V = 3 L
T = 350°C = 623 K
P = .990 atm
R = .0821 $\frac{L\cdot atm}{mol\cdot K}$
n = ?

$PV = nRT$

$(.990)(3) = n(.0821)(623)$

$n = .0581\ \cancel{moles}\ C_3H_8\left(\frac{4\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ C_3H_8}}\right)\left(\frac{18\,g}{1\ \cancel{mol\ H_2O}}\right) = \boxed{4.18\,g\ H_2O}$

44. To produce 15.4 L of nitrogen dioxide at 310 K and 2.0 atm, how many liters of nitrogen gas ($N_2$) and oxygen gas ($O_2$) are required? How many grams of each gas are required?

$PV = nRT$ P = 2 atm, T = 310 K, R = .0821

$$2N_2 + O_2 \longrightarrow 2N_2O$$

$15.4\,L\,N_2O\left(\frac{2\ mol\ N_2}{2\ mol\ N_2O}\right) = \boxed{15.4\,L\,N_2\ needed}$

$(2)(15.4) = n(.0821)(310)$

$n = 1.21\ mol\ N_2 \times 28\,g/mole = \boxed{33.9\,g\,N_2}$

$15.4\,L\,N_2O\left(\frac{1\ mol\ O_2}{2\ mol\ N_2O}\right) = \boxed{7.7\,L\,O_2\ needed}$

$(2)(7.7) = n(.0821)(310)$

$n = .605\ mol\ O_2 \times 32\,g/mole = \boxed{19.4\,g\,O_2}$