Key

# Unit 5 –Gas Laws

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## Objectives:

Physical Characteristics of Gases

- Use the kinetic molecular theory (KMT) to explain how certain physical properties of ideal gases differ from real gases
- Describe the conditions under which real gases deviate from "ideal" behavior
- Explain the five postulates of the kinetic molecular theory.
- Define pressure and standard pressure in terms of force and explain how pressure is measured
- Use the gas laws to express simple mathematical relationships between the pressure, temperature, volume and quantity of gases
- Relate KMT to the individual gas laws

Molecular Composition of Gases

- Develop a relationship between the volume, mass, and number of particles of a gas
- Perform calculations using the ideal gas law, specifically being able to solve for pressure, volume, temperature, or number of moles
- Perform stoichiometric calculations using the relationship between volume and moles both at STP and non STP conditions
- Demonstrate the relationship between the mass of gas particles and their rate of effusion

## Vocabulary:

- Kinetic molecular theory
- Gas pressure
- Gas temperature
- Ideal gas
- Ideal gas constant (R)
- Standard temperature and pressure (STP)
- atmosphere (atm)
- Kelvin (K)
- Pascal (Pa)
- Dalton's law of partial pressures
- Effusion
- Diffusion
- Graham's law
- Millimeter of mercury/Torr (mmHg)
- Kinetic energy
- Avogadro's Law
- Boyle's Law
- Charles' Law
- Gay-Lussac's Law
- rate

## Properties of Gases

To understand why a gas behaves as it does, we need to be able to picture what happens to gas particles as conditions (like temperature and pressure) change.

The Kinetic Molecular Theory (the theory of moving molecules) provides this picture. It is summarized below.

1. Gases consist of large numbers of molecules that are in constant, random motion.

2. The individual molecules of a gas essentially have zero volume compared to the total volume of the gas.

3. Attractive and repulsive forces between gases are negligible.

4. The particles move in straight lines.

5. The average energy of the particles is directly proportional to absolute temperature (Kelvin). At any given temperature all gases have the same kinetic energy.

If a gas behave in accordance with the above 5 postulates it is referred to as ideal. Real gases deviate from which of the postulates?

from postulate #2: Individual gas molecules of a "real" gas have a significant volume and mass (∴ are large molecules; the gas has a large molar mass)

from postulate #3: Molecules of a "real" gas have significant attractive/repulsive forces (∴ can undergo chemical reactions; tend to be polar)

Scientists use the "Gas Laws" to describe what will happen to gases under certain circumstances.

There are several gas laws that we will discuss, but they are all derived from the same basic gas equation:

$$PV = nRT$$

As you can see, this equation has many variables. Each one will have a different affect on how a gas behaves.

**The Variables:**

**1. Temperature (T)**

a. Defined as the average kinetic energy that the particles of the gas possess.

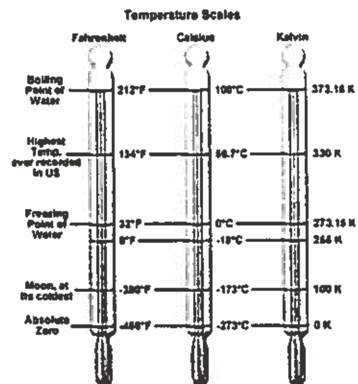
b. The number used to represent temperature depends on the scale being used. There are 3 scales:

   i. Fahrenheit, Celsius, and Kelvin.

      1. Kelvin is considered the absolute temperature scale because 0 K is called absolute 0 and is theoretically the lowest possible temperature and at this temperature the kinetic energy is zero.

c. Most thermometers we use record <u>Celsius and/or Fahrenheit</u>. Very few thermometers record Kelvin. To determine Kelvin we must do a conversion.

i. To convert temperature into Kelvin, you must start with a temperature in <u>Celsius!</u>




Use: $T_K = T_C + 273$

Example: 100°C = 373 K

ii. All temperature in gas law problems must be in Kelvin!

**2. Pressure (P)**

a. Defined as the force exerted on an object divided by the area over which the force is exerted.

$$pressure = \frac{force}{area}$$

i. Exerting more force increases the pressure
ii. Exerting a force over a smaller area increases the pressure.
iii. We use air under pressure (compressed air) for many everyday uses like filling our car tires with air and to fill SCUBA diving tanks.
iv. Pressure is the result of billions of gas particles moving and colliding against the container in which the gas is found.

b. Units of Pressure – There are many units used to measure pressure.

The SI unit for pressure is

| atm Atmospheres | mm Hg | torr | psi pounds per square inch | kPa kilopascals | Pa Pascals |
|---|---|---|---|---|---|
| 1 | 760 | 760 | 14.7 | 101.325 | $1.01325 \times 10^5$ |

c. Standard Temperature and Pressure – STP

i. This is the temperature and pressure which is universally known all around the world to be 273K and 101.325 kPa
ii. Sometimes STP uses the pressure unit "atmosphere" instead, in which case STP is 273K and 1 atm.

d. List three ways you can increase the pressure

i. Decrease volume: compress the gas
ii. Increase temperature: heat the gas
iii. Increase the # of moles: add gas to the container

3. **Volume (V)**
   a. Defined as the amount of space the gas takes up.
   b. Units of Volume – There are many units used to measure volume.

The SI unit for volume
$1\ dm^3$ = 1 liter as well as:

| Liter | $cm^3$ | quarts |
|---|---|---|
| 1 | 1000 | 1.057 |

   c. Remember that gas is the one phase of matter that does not have a definite volume. The gas laws we will be discussing will describe how/when the volume of a gas changes.

4. **Moles (n)**
   a. The term mole is used to describe a certain number of gas particles.
   b. Remember that 1 mole = $6.02 \times 10^{23}$ particles
   c. Avogadro discovered that 1 mole of any gas (provided it is at 1 atm of pressure and 273K) has exactly **$6.02 \times 10^{23}$ particles and 22.4 L = 1 mol.**
   d. In chemistry, this is useful because you can relate the number of particles in 1 mole of any gas to its atomic mass. For example:

1 mole oxygen = $6.02 \times 10^{23}$ particles = 32 grams= 22.4 Liters at STP

### *The Volume-Amount of Gas Relationship: Avogadro's Law*

- As the amount of gas increases in an elastic container, the volume will increase (pressure and temperature held constant.
- As the amount of gas decreases in an elastic container, the volume will decrease (pressure and temperature held constant.
- Why does this happen? Adding more gas molecules will require more volume to hold them in. More gas molecules need more space. (and vice versa).

Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

OR $V_1 n_2 = V_2 n_1$

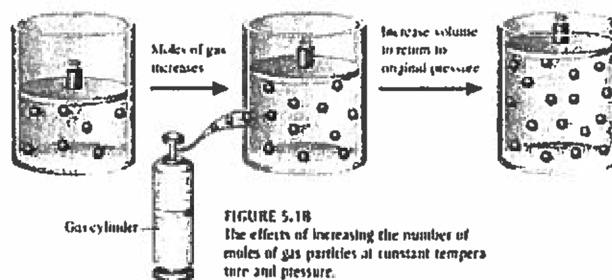


FIGURE 5.18 The effects of increasing the number of moles of gas particles at constant temperature and pressure.

**Sample Problem:**

1. 5.00 L of a gas is known to contain 0.965 mol. If the amount of gas is increased to 1.80 mol what new volume will result (at an unchanged temperature and pressure)?
   $V_1 n_2 = V_2 n_1$
   5.00 L(1.80 mol) = $V_2$(0.965 mol)
   $V_2$ = 9.33 L
2. 13.1 mol of a gas is in a 4.37 L container at constant pressure. The volume of the container decreases to 2.72 L how many moles of remain? $n_2$ = ?
   4.37 L ( $n_2$ ) = 2.72 L (13.1 mol)
   $n_2$ = 8.15 mol remains (in other words, 4.95 mol of gas escaped)

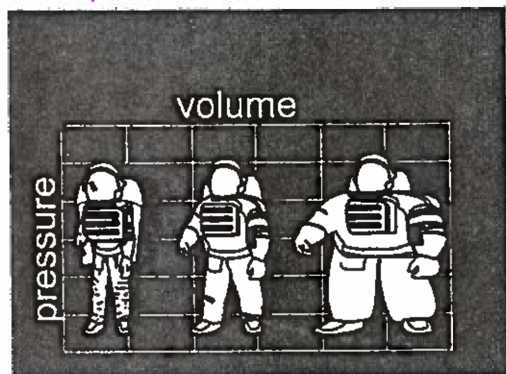
## The Pressure – Volume Relationship: Boyle's Law

- The first person to investigate the relationship between the pressure of a gas and its volume was the British chemist **Robert Boyle** (1627 – 1691).
- As the pressure of a gas increases, its volume decreases and as the pressure of a gas decreases, the volume increases (temperature held constant).
- As the volume of a gas increases, its pressure decreases and as the volume of a gas decreases, its pressure increases (temperature held constant).
- Why does this happen? As volume ↑, the space between gas molecules increase. There are now less chances for collision of the particles against the walls. (and vice versa)

Boyle's Law

$P_1V_1 = P_2V_2$

note: V and P units must be consistent. If different P and V units are used in a problem you'll have to convert.




**Sample Problem:**

1. The gas in a balloon has a volume of 4L when it is at 100kPa of pressure. The balloon is released into the atmosphere, where it expands to a volume of 8L. What is the new pressure on the balloon?

$P_1\ V_1\ = P_2\ V_2$

$(100\text{ kPa})(4\text{ L}) = (P_2)\ (8\text{ L})$

$\frac{100\text{kPa}\ (4\text{L})}{8\text{L}} = P_2$

$P_2 = 50\text{ kPa}$

2. Consider a 1.53 L sample of a gaseous $SO_2$ at a pressure of $5.6 \times 10^3$ Pa. If the pressure is changed to $1.5 \times 10^4$ Pa at constant temperature, what will be the new volume of gas? $V_2 = ?$

$P_1V_1 = P_2V_2$

$(5.6\times10^3\text{ Pa})(1.53\text{ L}) = (1.5\times10^4\text{ Pa})V_2$

$0.57\text{ L} = V_2$ (in other words... the gas volume decreases since the pressure ↑)

## The Temperature – Volume Relationship : Charles' Law

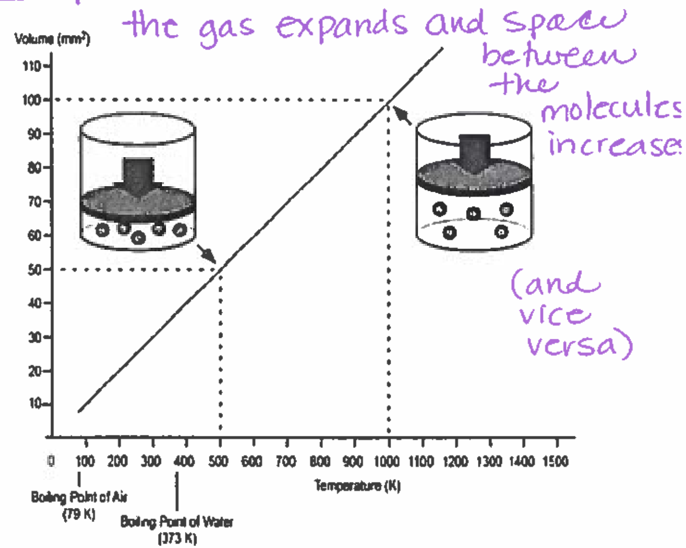
- **Jacques Charles** (1746-1823) was the scientist who developed the scientific law that relates temperature of a gas to its volume.
- As the temperature of a gas increases, its volume increases and as the temperature of a gas decreases, its volume decreases (pressure is held constant).
- As the volume of a gas increases, its temperature increases and as the volume of a gas decreases, its temperature decreases (pressure is held constant).

- Why does this happen? As gas molecules are warmed, their average kinetic energy increases. The molecules all move more and they will need more space to move around in, so the gas expands and space between the molecules increase (and vice versa)

Charles' Law

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

* must use Kelvin temps




**Sample Problem:**

1. Gas in a balloon occupies 2.5 L at 300K. At what temperature will the balloon expand to 7.5 L?

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

$$T_2 = \frac{7.5 \cdot 300}{2.5} \qquad T_2 = 900\,K$$

2. A gas with a volume of 600. mL has a temperature of 30°C (→ 303 K). At constant pressure the gas is heated until the gas expands to 1,200 mL. What is the new temperature of the gas if the pressure remains constant?

→ If $T_1$ is in °C then express $T_2$ in °C as well

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{so} \quad \frac{600.\,mL}{303\,K} = \frac{1200\,mL}{T_2}$$

(cross multiply)

$600.\,mL\,(T_2) = (1200\,mL)(303\,K)$

$T_2 = 610\,K$ (2 sig figs)

or The volume doubles – so does the Kelvin temperature

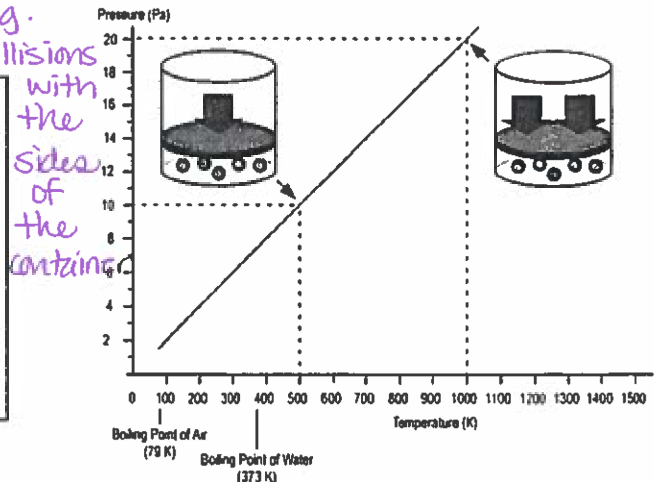
**The Temp. – Pressure Relationship: Gay-Lussac's Law**

- **Joseph Gay-Lussac** explored the relationship between the temperature of a gas and its pressure.
- As the pressure of a gas increases, its temperature increases and as the pressure of a gas decreases, its temperature decreases (volume held constant).
- As the temperature of a gas increases, its pressure increases and as the temperature of a gas decreases, its pressure decreases (volume held constant).
- Why does this happen? when T ↑, avg. kinetic energy ↑ too. This ↑ # of collisions with the sides of the container

Gay-Lussac's Law

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

* must use Kelvin temps



**Sample Problem:**

1. The pressure of a gas in a tank is 4.20 atm at 44°C. If the temperature rises to 67°C, what will be the gas pressure in the tank?

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

$$= \frac{4.20\,atm\,(340\,K)}{317}$$

$$P_2 = 4.50\,atm$$

2. Ten liters of a gas is found to exert 97.0 kPa at 25.0°C. What would be the required temperature (in Celsius) to change the pressure to standard pressure?

→ 298 K  $T_2 = ?$

→ 101.325 kPa $= P_2$

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

$$\frac{97.0\,kPa}{298\,K} = \frac{101.325\,kPa}{T_2} \quad \text{cross multiply}$$

$$97.0\,kPa\,(T_2) = (101.325\,kPa)(298\,K)$$

$T_2 = 311$ K ; convert so $T_2 = 38.0°C$

**Combined Gas Law**

There are times when temperature, volume, and pressure are all affected when the conditions of a gas change. In these instances, we must combine Boyle's Law, Charles' Law, and Gay-Lussac's Law into what is known as the **Combined Gas Law.** Knowing the combined gas laws gives the individual gas laws by holding one of the variables constant.

> Combined Gas Law
>
> $$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$
>
> * use if P, V, and T all change

**Sample Problem:** $P_1$ ?

Boyle's Law

1. What pressure will be needed to reduce the volume of 77.4 L of helium at 98.0 kPa to a volume of 60.0 L?

$$P_1 V_1 = P_2 V_2$$

$$(98.0\,kPa)(77.4\,L) = P_2\,(60.0\,L)$$

$$126\,kPa = P_2$$

(expect $P_2$ to be higher than $P_1$)

Boyle's Law

2. At 250.0 mL sample of chlorine gas is collected when the barometric pressure is 105.2 kPa. What is the volume of the sample after the barometer drops to 100.3 kPa? $V_2 = ?$

$$P_1 V_1 = P_2 V_2$$

$$(105.2\,kPa)(250.0\,mL) = (100.3\,kPa)V_2$$

$$262.2\,mL = V_2$$

(expect $V_2$ to be higher since the P drops)

Charles' Law

3. A sample of $SO_2$ gas has a volume of 1.16 L at a temperature of 23°C. At what temperature will the gas have a volume of 1.25 L? $T_2 = ?$ → 296 K

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

$$\frac{1.16\,L}{296\,K} = \frac{1.25\,L}{T_2}$$

$$1.16\,L\,(T_2) = (1.25\,L)(296\,K)$$

$T_2 = 319$ K ; convert so $T_2 = 46°C$.

(If the gas volume ↑, T should ↑ too)

4. The balloon is released from a 4000 m mountaintop where the pressure is 61.7 kPa. What is the volume of the balloon when it is released?

5. A balloon is inflated with 6.22 L of helium at a temperature of 36°C. What is the volume of the balloon when the temperature is 22°C? → 295 K

→ 309 K; → $V_2$?

Charles' Law

(If T↓ the V should ↓ too)

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

$$\frac{6.22\ L}{309\ K} = \frac{V_2}{295\ K}\ ;\ 309\ K(V_2) = 6.22\ L(295\ K)$$

$V_2 = 5.94\ L$

6. Helium gas in a 2.00 L cylinder is under 1.12 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature of the gas in the cylinder? $T_1 = ?$

Gay-Lussac's Law

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

$$\frac{1.12\ atm}{T_1} = \frac{2.56\ atm}{309.5\ K}\ ;\ 2.56\ atm\ (T_1) = (1.12\ atm)(309.5\ K)$$

$T_1 = 135\ K$ so in °C, $T_1 = -138$

7. If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how much does the temperature have to decrease to lower the pressure to 28.4 kPa?

→ $T_2$? then subtract $T_1$

Gay-Lussac's Law

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

$$\frac{30.7\ kPa}{273\ K} = \frac{28.4\ kPa}{T_2}\ ;\ 30.7\ kPa\ (T_2) = (28.4\ kPa)(273\ K)$$

$T_2 = 253\ K$ ; so T must decrease by 20 K

8. A sample of ammonia gas occupies a volume of 1.58 L at 22°C and a pressure of 0.983 atm. What volume will the sample occupy at 1.00 atm and 0°C? $V_2 = ?$

Combined Gas Law

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

$$\frac{(0.983\ atm)(1.58\ L)}{295\ K} = \frac{(1.00\ atm)V_2}{273\ K}\ ;\ (295\ K)(1.00\ atm)V_2 = (0.983)(1.58)(273)$$

$V_2 = 1.44\ L$

9. A student collects 285 mL of $O_2$ gas at a temperature of 15°C and a pressure of 99.3 kPa. The next day, the same sample occupies 292 mL at a temperature of 11°C. What is the new pressure of the gas? $P_2$?

Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}\ ;\ \frac{(99.3\ kPa)(285\ mL)}{288\ K} = \frac{P_2(292\ mL)}{284\ K}$$

$$(288\ K)(P_2)(292\ mL) = (99.3\ kPa)(285\ mL)(284\ K)$$

$P_2 = 95.5\ kPa$

10. You have a 2.40 L container of air at STP. From out of nowhere, Bigfoot stomps on it, decreasing the container's volume down to 0.500 L and increasing the pressure to 8.00 atmospheres. What is the temperature of the air in the container? $T_2$?

STP = 1 atm and 0°C

Combined Gas Law

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

$$\frac{(1\ atm)(2.40\ L)}{(273\ K)} = \frac{(8.00\ atm)(0.500\ L)}{T_2}\ ;\ (1\ atm)(2.40\ L)T_2 = (8.00)(0.500)(273)$$

$T_2 = 455\ K = 182\ °C$

11. A 12.0 L sample of $NO_2$ gas is at STP. What would be its new volume if its pressure was decreased to 575 mmHg and its temperature was doubled? $V_2$?

STP = 760 mmHg and 0°C

Combined Gas Law

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

$$\frac{(760\ mmHg)(12.0\ L)}{(273\ K)} = \frac{(575\ mmHg)V_2}{546\ K}$$

$$(760\ mmHg)(12.0\ L)(546\ K) = (273\ K)(575\ mmHg)V_2$$

$31.7\ L = V_2$

$V_2 = ?$

12. A sample of Cl2 gas occupies a volume of 11.4 L at 3.50 atmospheres. When the Cl2 is changed to STP conditions, what will be its volume? Ti not given so you can assume T is constant

Boyle's Law

$P_1V_1 = P_2V_2$

$(3.50 \text{ atm})(11.4 \text{ L}) = (1 \text{ atm}) V_2$

$39.9 \text{ L} = V_2$

13. You fill your car's tires to 35 psi when they were cold (12°C). After driving for 3 hours, your car's tires warm up to 38°C. What would be the pressure inside your tires now, in psi? $P_2 = ?$

Gay-Lussac's Law

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

$\frac{35 \text{ psi}}{285 \text{ K}} = \frac{P_2}{311 \text{ K}}$ ; $285 \text{ K} (P_2) = 35 \text{ psi} (311 \text{ K})$

$P_2 = 38.1 \text{ psi}$

14. You're at the zoo and have a big red 1.80 L helium balloon. The barometric pressure today is 785 mmHg. Then you hear the roar of a lion. Startled, you accidentally release the balloon. It flies away. By the time it reaches the clouds, the atmospheric pressure that high is only 0.300 atmospheres. What would be the volume of the balloon up there? $V_2 = ?$

Boyle's Law

$\frac{0.300 \text{ atm} \mid 760 \text{ mmHg}}{\mid 1 \text{ atm}} = 288 \text{ mmHg}$

$P_1V_1 = P_2V_2$

$(785 \text{ mmHg})(1.80 \text{ L}) = (288 \text{ mmHg}) V_2$

$\frac{(785 \text{ mmHg})(1.80 \text{ L})}{(288 \text{ mmHg})} = 4.90 \text{ L} = V_2$

15. A container of oxygen gas is at STP. If this sample is put into an oven at 280°C, what would its pressure be, in atmospheres? $P_2 = ?$

Gay-Lussac's Law

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

$\frac{1.00 \text{ atm}}{273 \text{ K}} = \frac{P_2}{553 \text{ K}}$ ; $273 \text{ K} (P_2) = (1.00 \text{ atm})(553 \text{ K})$

$P_2 = 2.0 \text{ atm}$

## Ideal Gas Law

**PV = nRT**

Volume, pressure, temperature, and amount of gas can be compared by using a conversion factor known as the **ideal gas constant (R)**. While the term constant implies not changing, the value of the gas constant is dependent on the unit used to measure the pressure.

Gas Constant Values

| If the Units of pressure is: | The Numerical Value of R Is (with units): |
|---|---|
| atm | 0.0821 L-atm/mole-K |
| kPa | 8.314 L-kPa/mole-K |
| mmHg (Torr) | 62.4 L-Torr/mole-K |

**PV = nRT** **$R = \frac{PV}{nT}$** **To solve in units of pressure: atm** **$\frac{1\text{atm}(22.4\text{L})}{1 \text{ mol}(273\text{K})} = 0.0821 \frac{\text{atm.L}}{\text{mol.K}}$**

***THE CONSTANT (R) THAT IS USED IS BASED ON THE UNITS OF PRESSURE**

1. If I have 4 moles of a gas at a pressure of 5.6 atm and a volume of 12 liters, what is the temperature?

2. If I have an unknown quantity of gas at a pressure of 120 kPa, a volume of 31 liters, and a temperature of 87 $^{0}$C, how many moles of gas do I have?

3. If I contain 3 moles of gas in a container with a volume of 60 liters and at a temperature of 400 K, what is the pressure in mmHg inside the container?

4. If I have 7.7 moles of gas at a pressure of 0.09 atm and at a temperature of 56 $^{0}$C, what is the volume of the container that the gas is in?

## The Ideal and Combined Gas Laws

$$PV = nRT \quad \text{or} \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Use your knowledge of the ideal and combined gas laws to solve the following problems. If it involves moles or grams, it must be PV = nRT

1. If four moles of a gas at a pressure of 5.4 atmospheres have a volume of 120 liters, what is the temperature?

2. If I initially have a gas with a pressure of 84 kPa and a temperature of 35°C and I heat it an additional 230 degrees, what will the new pressure be? Assume the volume of the container is constant.

1. If I have 4 moles of a gas at a pressure of 5.6 atm and a volume of 12 liters, what is the temperature?

$PV = nRT$  $T = 204.6\,^{\circ}K$

$(5.6)(12) = 4\,(.0821)\,T$

2. If I have an unknown quantity of gas at a pressure of 120 kPa, a volume of 31 liters, and a temperature of 87 $^{0}$C, how many moles of gas do I have?

$PV = nRT$  $n = 1.24$ moles

$(120)(31) = n(8.314)(87+273)$

3. If I contain 3 moles of gas in a container with a volume of 60 liters and at a temperature of 400 K, what is the pressure in mmHg inside the container?

$PV = nRT$  $P = 1248$ mm

$P(60) = 3\,(62.4)(400)$

4. If I have 7.7 moles of gas at a pressure of 0.09 atm and at a temperature of 56 $^{0}$C, what is the volume of the container that the gas is in?

$PV = nRT$  $(.09)(V) = (7.7)(0.0821)(56+273)$

$V = 2310$ L

## The Ideal and Combined Gas Laws  PV = nRT or $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Use your knowledge of the ideal and combined gas laws to solve the following problems. If it involves moles or grams, it must be PV = nRT

1. If four moles of a gas at a pressure of 5.4 atmospheres have a volume of 120 liters, what is the temperature?

$PV = nRT$

$(5.4)(120) = (4)(0.0821)\,T$

$T = 1973\,^{\circ}K$

2. If I initially have a gas with a pressure of 84 kPa and a temperature of 35°C and I heat it an additional 230 degrees, what will the new pressure be? Assume the volume of the container is constant.

$P = 146.7$ kPa

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$  $\frac{84}{35+273} = \frac{P}{35+273+230}$

3. My car has an internal volume of 2600 liters. If the sun heats my car from a temperature of 20°C to a temperature of 55°C, what will the pressure inside my car be? Assume the pressure was initially 760 mm Hg.

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ $\frac{(2600)(760)}{20+273} = \frac{?(2600)}{55+273}$

P = 850.78 mm

4. How many moles of gas are in my car in problem #3?

PV = nRT

(760)(2600) = n (62.4)(20+273)

n = ~~0.0093~~ 108.1

5. A toy balloon filled with air has an internal pressure of 1.25 atm and a volume of 2.50 L. If I take the balloon to the bottom of the ocean where the pressure is 95 atmospheres, what will the new volume of the balloon be? How many moles of gas does the balloon hold? (Assume T = 285 K)

$P_1V_1 = P_2V_2$ (1.25)(2.5) = (95)(V)

V = 0.033 L

PV = nRT

(1.25)(2.5) = n(0.082)(285)

n = 0.134

## Molar Volume and Reactions of Gases

As you solve the following problems keep in mind Avogadro's law, which states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules. From Avogadro's law, it follows that all gases have equal molar volumes if they are measured at the same temperature and pressure. The molar volume is the volume occupied by one mole of a substance. At standard temperature and pressure (STP—0°C and 1 atm), the molar volume of any gas sample is 22.4 L.

1. What volume would be occupied by 2.0 mol nitrogen, $N_2$, gas at 0°C and 1 atm? 2 × 22.4

Volume = 44.8 L

2. What volume would be occupied by 88.0 g of gaseous carbon monoxide, CO, at STP?

12+16 = 44

Volume = 44.8 L

2 mole CO

## Gas Stoichiometry Practice

The amount of gas can be determined from the amount of another substance using stoichiometery. If the reaction takes place at STP, the molar volume relationship of 1 mol occupies 22.4 L can be used, however, if the reaction takes place at conditions other than STP, the ideal gas equation must be used to solve for the mols of gas.

1. Tanks of gaseous propane are used for cooking and heating. When propane ($C_3H_8$) burns (using oxygen from the air), the products of the reaction are carbon dioxide and water vapor.

   a. Write a balanced equation for this reaction.


   b. At constant temperature and pressure, how many liters of oxygen would be needed to completely combust 0.350 L of propane? Assum STP 22.4 L/m

.0156 m Prop $\frac{5}{1}$ = .078 m $O_2$

1.75 L $O_2$

Volume = 1.75 L $O_2$

   c. Continuing from 3b, how many liters of water vapor would be produced by the reaction of 0.350 L of propane?

1:4 0.0624 m $H_2O$ × 22.4

Volume = 1.39 L $H_2O$

2. Hydrogen chloride gas can be produced by a reaction between hydrogen gas and chlorine gas.

   a. Write a balanced equation for this reaction.


   b. At constant temperature and pressure, how many liters of hydrogen are needed to produce 1.75 L of hydrogen chloride? STP 22.4 L/m

1.75 L ⇒ 0.078 m HCl 2:1

$H_2$ 0.039 m $H_2$

Volume = 0.875 L $H_2O$

   c. Continuing from 4b, how many moles of chlorine would be needed to react with 8.65 mol of hydrogen?

8.65 mole $Cl_2$ 1:1

Volume = 193.76 L $Cl_2$

3. The industrial production of ammonia proceeds according to the following Equation.


   a. If 20.0 mol of nitrogen is available, what volume of $NH_3$ at STP can be produced?

20 $\frac{2}{1}$ = 40 $NH_3$ × 22.4

896 L $NH_3$

   b. What volume of $H_2$ at STP will be needed to produce 800. L of ammonia, measured at 55°C and 0.900 atm?

PV = nRT

(.9)(800) = n (0.0821)(55+273)

$NH_3$ n = 26.73 $\frac{3\ H_2}{2\ NH_3}$

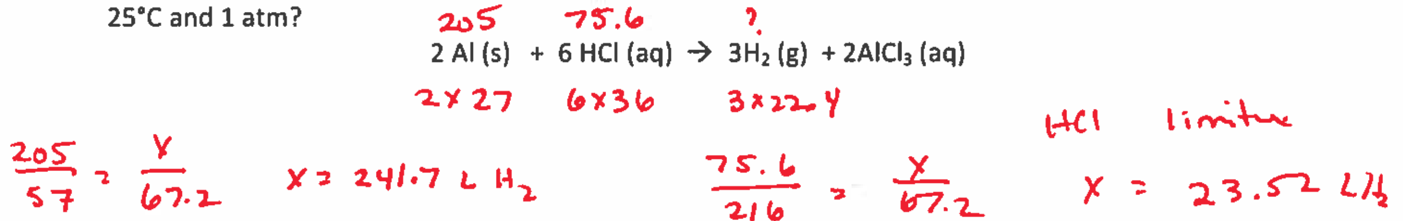
898.3 L $H_2$

40.1 m $H_2$ × 22.4 STP

**Limiting Reactant Stoichiometry**

Sometimes during a reaction it is necessary to make sure the reaction goes to completion, meaning one of the reactants is used up completely and the other reactants are left in excess. The reactant that is used completely is the **limiting reactant**. Once all of the limiting reactant is used the reaction will stop.

1. If you start with 205 g of aluminum and 75.6 g of HCl, a. how many liters of $H_2$ can be produced at 25°C and 1 atm?

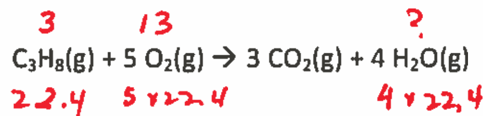


$$2\ Al\ (s)\ +\ 6\ HCl\ (aq) \rightarrow 3H_2\ (g)\ +\ 2AlCl_3\ (aq)$$

b. How many grams of excess reactant are left-over?

c. Suppose you only made 1.98 g of H2. What is your % yield?

2. Propane burns according to the following equation.

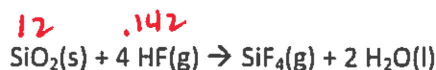


$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

a. What volume of water vapor measured at 250°C and 1.00 atm is produced when 3.0 L of propane and 13.0 L of $O_2$ at STP are burned?

b. What volume of oxygen at 20.°C and 102.6 kPa is used if 640. L of $CO_2$ is produced? The $CO_2$ is also measured at 20.°C and 102.6 kPa.

3. Silicon tetrafluoride gas can be produced by the action of HF on silica according to the following equation.




60.g/m

1.00 L of HF gas under pressure at 3.48 atm and a temperature of 25°C reacts with 12.0 g of $SiO_2$ to form $SiF_4$. What volume of $SiF_4$, measured at 15°C and 0.940 atm, is produced by this reaction?

HF $\frac{V}{n} = \frac{RT}{P} = \frac{(0.0821)(298)}{3.48}$ $SiF_4$ $\frac{V}{n} = \frac{RT}{P} = \frac{(0.0821)(15+273)}{0.940} = 25.15$ L/m

n HF = 0.142 $\frac{12}{60} = \frac{x}{25.15}$ 5.03 L $SiF_4$ 12 g $SiO_2$

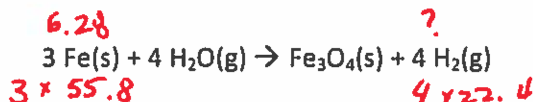
.142 HF $\frac{1}{4}$

.036 m $SiF_4$

→ 0.89 L

Limiting HF

4. One method used in the eighteenth century to generate hydrogen was to pass steam through red-hot steel tubes. The following reaction takes place.




a. What volume of hydrogen at STP can be produced by the reaction of 6.28 g of iron?

$\frac{6.28}{167.4} = \frac{x}{89.6}$ x = 3.36 L $H_2$

b. What mass of iron will react with 500. L of steam at 250.°C and 1.00 atm pressure?

$n_{H_2O} = \frac{PV}{RT} = \frac{(1)(500)}{(0.0821)(250+273)} = 11.64$ moles $H_2O$ $\frac{3}{4} = 8.73$ mole Fe

8.73 × 55.8 = 487.3 g Fe

c. If 285 g of$Fe_3O_4$ are formed, what volume of hydrogen, measured at 20.°C and 1.06 atm, is produced?

→ 231.4 g/m ⇒ 1.23 m $Fe_3O_4$ $\frac{4}{1}$ = 4.93 mole $H_2$

$V = \frac{nRT}{P} = \frac{(4.93)(0.0821)(20+273)}{1.06} = 111.8$ L $H_2O$

5. It is possible to generate chlorine gas by dripping concentrated HCl solution onto solid potassium permanganate according to the following equation.


If excess HCl is dripped onto 15.0 g of $KMnO_4$, what volume of $Cl_2$ will be produced? The $Cl_2$ is measured at 15°C and 0.959 atm.

→ 158 g/m = 0.095 m $KMnO_4$ $\frac{5}{1}$ = 0.475 m $Cl_2$

$V = \frac{nRT}{P} = \frac{(0.475)(0.0821)(15+273)}{0.959} = 11.7$ L $Cl_2$

**Dalton's Law of Partial Pressure:** Dalton stated that the sum of the pressures of individual gases in the same container is equal to the pressure of the container. $P_{tot}=P_1+P_2+P_3...$

### Effusion, Diffusion, and Graham's Law

Gases are a fluid substance in constant random motion. It is possible to measure the velocity of a gas based on its kinetic energy. Before beginning let's review a few concepts.

- What exactly is temperature a measurement of? Average K.E
- Why is it important to include the word "average" in your answer? Each molecule independ motion
- What two factors does an object's kinetic energy depend on? mass and velocity
- What specifically is the equation for kinetic energy? $KE = \frac{1}{2}mv^2$
- Which would increase the kinetic energy of an object more: doubling the object's mass or doubling the objects velocity? velocity Explain: $\frac{1}{2}mV^2$
- Define diffusion rate of gas mixing
- Define effusion " " escape thru pin hole

State Graham's Law as an equation for two gases (A and B) at the same temp:

$$\frac{V_a}{V_b} = \sqrt{\frac{m_b}{m_a}}$$

Based on Graham's Law the velocity of a gas varies inversely with its mass in kilograms.

**Consider two gases, He and O2, at the same temperature...** $He = 4$ $O_2 = 32$

- Which particles would have greater average kinetic energy? He at same Temp.
- Which particles are heavier? $O_2$
- Which particles would have greater velocity? He
- Which gas would diffuse across the room faster? He

**Two gas samples, one H2 and one CO2, are such that their particles have the same velocity...** $m_{CO_2} > m_{H_2}$

- Which gas molecules have the greater average kinetic energy? $CO_2 > H_2$
- Which gas is at the higher temperature? $CO_2$ Explain: $\frac{1}{2}mv^2 = \frac{3}{2}kT$

***For the following questions, use the Graham's Law equation. Show all work.***

- At a certain temperature, $O_2$ molecules move with an average velocity of 345 mph. At that same temperature, what would be the average velocity of **a)** He atoms? **b)** $CO_2$ molecules?

  4 44

  $$\sqrt{\frac{m_{O_2}}{m_{He}}} = \frac{V_{He}}{V_{O_2}}$$

  Ans: **a)** 976 **b)** 294

- At a certain temperature, $CH_4$ molecules move with an average velocity of 187 m/sec. At that same temp, gas X particles have an average velocity of 141 m/sec. **a)** Is gas X heavier or lighter than $CH_4$? **b)** What is the molecular weight of gas X? **c)** What is a possible identity of gas X?

  16

  $$\sqrt{\frac{16}{x}} = \frac{141}{187}$$

  $x = 28.14$

  Ans: **a)** heavier **b)** 28.14 **c)** $N_2$