CHAPTER 14
Gases

14.1 The Gas Laws

Gas Particles and the Kinetic Theory?

> Gases do not attract or repel
> Gas particles are in constant motion
> Gas particles have elastic collisions
> They have the same average kinetic energy a given Temperature

Volume of Gases

- **Compressibility**:the volume of matter decreasing under pressure.
- Gases are easily compressed due to the large amount of space between gas particles.
- The actual volume of the particles is far less than the volume they take up.

Factors affecting Gas Pressure

1. amount of gas *(Moles)*
2. volume (Liters)
3. temperature (Kelvin)

Effect of adding or removing a Gas

- Increasing or decreasing the amount
- Pressure caused by collisions
- Addition of gas increases collisions
- Direct relationship

Add Particles = increase pressure

Remove Particles = decrease pressure

Effect of heating or cooling a Gas

- Raising the temperature, increases the pressure
- Collisions and kinetic energy
- Temperature in Kelvin

Effect of Changing the Container Size

Reduce volume. . . increase pressure

Increase volume...decrease pressure

Inverse relationship

14.1 Key Concept

In a closed container, give 3 ways you can increase the pressure by 5?

1. **Increase the amount of gas**
2. **Increase the temperature**
3. **Decrease the volume**

Pressure Unit Conversions

101.3 kPa
1 atm
760 mmHg
760 Torr

Equal Values

4 Factors that can change Gases

1. Mass
2. Volume
3. Temperature
4. Pressure

Boyle's Law

- For a given mass of gas at constant temperature, the volume of the gas varies inversely with pressure.
- Mass & Temperature must remain constant
- Requires a flexible container

Initial — Final

$$P_1 \times V_1 = P_2 \times V_2$$

Relationship of Volume and Pressure

If volume doubles, the pressure is halved.

- Why?

If volume is cut in half, the pressure is doubled.

- Why?

Boyle's Law Example Problem

A balloon contains 30.0 liters of a gas at 100.0 kPa. If mass and temperature are constant, what did the volume change to that caused the pressure to decrease to 25 kpa?

Constant
- Mass
- Temperature

P1:
V1:
P2:
V2:

What type of container is it?

Some gas laws cannot be used based on the type or container.

Rigid containers do not expand or contract.

- Their volume never changes! *(Constant Volume)*

Flexible containers expand or contract.

- Change based on the pressure inside. *(Volume not constant)*

Charles' Law

An increase in temperature causes a volume increase

An decrease in temperature causes a volume decrease

Direct Relationship in a flexible container

Pressure and mass remain constant

As temperature increases, particles move faster causing more collisions. To maintain constant pressure, the volume must increase.

Charles' Law

When using Charles Law, the container must be flexible! If it is not flexible, the pressure will not remain constant.

Kelvin Temperature must be used!

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

Charles' Law Example Problem

A balloon at 27 °C has a volume of 4.0 liters. Assuming the pressure and mass remain constant, what happens to the volume of the gas when it is heated to 57 °C?

Constant
- Mass
- Pressure

V1:
T1:
V2:
T2:

Gay-Lussac's Law

Temperature of a gas increases, pressure increases.

Temperature of a gas decreases, pressure decreases.

Direct relationship in a rigid container

Mass and Volume must remain constant

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

Gay Lussac's Law Example

The gas in an aerosol can has a pressure of 100 kPa at a temperature of 27 °C. Assuming volume and mass remain constant, what is the new pressure if the temperature is raised to 927 °C?

Constant

- Mass
- Volume

P1:

T1:

P2:

T2:

14.2 Combined Gas Law

The three gas laws are combined into a single expression.

> *Temperature & Mass constant : Boyle's*
> *Volume & Mass constant : Gay-Lussac's*
> *Pressure & Mass constant: Charles'*

Mass is the only thing constant.

Requires a flexible container

Combined Gas Law

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

$$P1 \times V1 \times T2 = P2 \times V2 \times T1$$

Combined Gas Law Example

A balloon has a volume of 30.0 liters, a pressure of 150 kPa, and a temperature of 40 °C. Assuming mass is constant, what is the new volume of the gas at STP? *(Hint: what is STP?)*

Constant

- Mass

P1:

V1:

T1:

P2

V2:

T2:

1. Temp. Constant/Mass Constant

If volume decreases then pressure increases.
If volume increases then pressure decreases.

2. Volume Constant/Temp. Constant

If mass of gas decreases then pressure decreases.
If mass of gas increases then pressure increases.

3. Volume Constant/Mass Constant

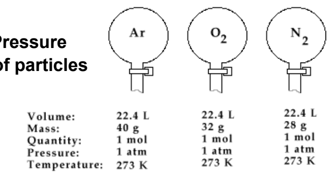
If temperature of gas decreases then pressure decreases.
If temperature of gas increases then pressure increases.

4. Pressure Constant/Mass Constant

If temperature of gas decreases then volume decreases.
If temperature of gas increases then volume increases.

14.2 Avogadro's Principle

1. **Equal Volumes of Gas**
2. **Same Temperature and Pressure**
3. **Contain equal numbers of particles**



14.3 Ideal Gas Law

- Brings in all 4 factors that impact gases.
  - > Mass, volume, temperature, & pressure
- Nothing constant
- A gases properties at one moment in time
- Nothing changes

P x V = n x R x T

Ideal Gas Constant

The ideal gas constant (R) is .0821 when using atm.

The ideal gas constant (R) is 8.31 when using kpa.

The ideal gas constant (R) is 62.4 when using torr.

Units of Ideal Gas Constant

The units of the ideal gas constant are used to cancel every unit except for your unknown.

$$\frac{\text{L x atm}}{\text{Mol x K}}$$

Rules for using Ideal Gas Constant

n = moles (convert if given grams)

V = must be in liters

T = must be in kelvin

P = kpa when using 8.31

P = atm when using .0821

Idea Gas Law - Example 1

A rigid steel cylinder with a volume of 20.0 L of N_2 gas reaches a final pressure of 20,000 kPa at 27 °C. How many moles of N_2 gas does the cylinder contain?

Ideal Gas Law - Example 2

You are given 27 grams of CH_4 in a container that is 5500 mL at a temperature of 56 °C. What is the pressure of the gas in the container?

Real vs. Ideal Gases

- **Ideal Gas**: have very small or no volume, no intermolecular attractions, and no elastic collisions
- **Real Gas**: when ideal gas law does not apply
  - > Gases at low temp and high pressures.
  - > Large gas molecules have too much volume
  - > polar molecules have too many attractions

Think, is it possible for gas particles to have "0" volume and "0" collisions?

Molar Mass & the Ideal Gas Law

Recall: $n = \frac{\text{mass } (m)}{\text{molar mass (M)}}$ $PV = \mathbf{nRT}$

Hence: $PV = \frac{m\text{RT}}{\text{M}}$ or $\text{M} = \frac{m\text{RT}}{\text{PV}}$

Density & the Ideal Gas Law

Recall: $D = \frac{\text{mass } (m)}{\text{volume (V)}}$

Hence: $\mathbf{D} = \frac{m}{\text{V}} = \frac{\text{MP}}{\text{RT}}$ or $\text{M} = \frac{\text{DRT}}{\text{P}}$

Ideal Gas Law - Example 3

An unknown gas @ STP has a density of 1.78 g/L. Determine the unknown gas by calculating the molar mass.

Ideal Gas Law Example 4

A sample of N_2 gas has a pressure of 3.51-atm and a temperature of 327 kelvin . What is the density ?

14.3 : Mixtures of Gases

Dalton's Law:
- used for a mixture of gases
- The contribution each gas in a mixture makes to the total pressure
- The pressure each gas contributes to the total pressure is the partial pressure.

Partial Pressures

- The proportionate pressure of each gas in a mixture does not change as Temperature, Volume, or total pressure changes.
- Mass is the only thing that can change the partial pressure.
- Although, the partial pressures do change as temperature, pressure, and volume change.

Total Pressure 1.0 atm	Proportion	Total Pressure 5.0 atm
N_2 = 0.65 atm		
O_2 = 0.20 atm		
CO_2 = 0.11 atm		
H_2 = 0.04 atm		

Partial Pressures Example 1

Air contains O_2, N_2, CO_2, and other gases. If the partial pressure of N_2 =79.10 kPa, CO_2 = 0.040 kPa, and other gases = 0.94 kPa. At STP, what is the partial pressure of O_2? What is the percentage of each gas in the air?

Nitrogen: 79.10 kPa
CO_2: 0.040 kPa
Other: 0.94 kPa

80.08 kPa

Total Pressure @ STP: ___________

Partial Pressures Example 2

An air sample contains 2.1 mol O_2, 1.5 mol N_2, and 4.0 mol CO_2 at a temperature of 47 °C in a container with a volume of 8.3 L. What is the total pressure of the air sample and what are the partial pressures of each gas?

O_2: $P = \frac{nRT}{V} = \frac{2.1 \text{ mol} \times .0821 \times 320 \text{ K}}{8.3\text{-L}}$

N_2: $P = \frac{nRT}{V} = \frac{1.5 \text{ mol} \times .0821 \times 320 \text{ K}}{8.3\text{-L}}$

CO_2: $P = \frac{nRT}{V} = \frac{4.0 \text{ mol} \times .0821 \times 320 \text{ K}}{8.3\text{-L}}$

Partial Pressures of Each Gas

Total Pressure = ___________

14.4 Gas Stoichiometry

- Ratios from balanced chemical equation can be used at constant pressure.
- When gases are mixing, temperature and pressure are not factors since they are in the same container.

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

At constant Temperature and pressure, how many liters of C_3H_8 will undergo complete combustion with 16.0-L of Oxygen gas?

16.0-L O_2	

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

Assuming STP, how many liters of C_3H_8 will undergo complete combustion with 34.0-L of Oxygen gas?

34.0-L O_2			

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

Assuming STP, how many grams of CO_2 are produced from 34.0-L of Oxygen gas?

34.0-L O_2			

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

If 6.50-L of nitrogen reacts completely at a constant pressure of 2.5-atm and a constant temperature of 308 K, how many grams of ammonia are produced?

6.5-L N_2	

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