GASES

AP CHEMISTRY 20

CHARACTERISTICS OF GASES

TOPIC ONE

CHARACTERISTICS OF GASES

- While many gases have different ***chemical properties*** they have similar ***physical properties***.
  - All gases will expand and occupy the entire of the volume of the container and can easily be compressed. It does not matter the identity of the gas (e.g., nitrogen gas, water vapor) each gas will behave similarly.
- Substances that can exist in the gaseous state but are usually in another state are said to be **vapours** (e.g., water vapour).

CHARACTERISTICS OF GASES

- Unlike most solutions, **gases form homogeneous mixtures**. When gasoline vapour and water vapor mix it is homogeneous while water and gasoline in the liquid form will not mix (forms an aqueous and organic layer).
- Gases are substances with **large distances between the molecules** (which allows for the compression) and most behave as though the other molecules are not present.
  - Only about 0.1% of any volume of a gas has particles present at any given moment.

TABLE 10.1 • Some Common Compounds That Are Gases at Room Temperature

Formula	Name	Characteristics
HCN	Hydrogen cyanide	Very toxic, slight odor of bitter almonds
H_2S	Hydrogen sulfide	Very toxic, odor of rotten eggs
CO	Carbon monoxide	Toxic, colorless, odorless
CO_2	Carbon dioxide	Colorless, odorless
CH_4	Methane	Colorless, odorless, flammable
C_2H_4	Ethene (Ethylene)	Colorless, ripens fruit
C_3H_8	Propane	Colorless, odorless, bottled gas
N_2O	Nitrous oxide	Colorless, sweet odor, laughing gas
NO_2	Nitrogen dioxide	Toxic, red-brown, irritating odor
NH_3	Ammonia	Colorless, pungent odor
SO_2	Sulfur dioxide	Colorless, irritating odor

PRESSURE

TOPIC TWO

PRESSURE

- Pressure is a type of force that tends to move something in a specific direction. Pressure, P, is defined by the force, F, that is exerted on an area, A.

$$P = \frac{F}{A}$$

- A balloon remains inflated due to the pressure the molecules exert on the edge of the balloon.

ATMOSPHERIC PRESSURE

- The atmosphere is filled with gases that are attracted to Earth due to the gravitational pull.
- While the mass of any single particle is not much, the mass of the atmosphere does exert a force on the objects (e.g., the surface of Earth, people).
- This force is evident when the atmosphere is removed (vacuum created) from a water bottle – it begins to be crushed due to the imbalance of internal and external forces.

ATMOSPHERIC PRESSURE

- The atmospheric pressure is determined by the force exerted on a 1 m^2 due to the *entire* atmosphere that resides above it – roughly 10,000 kg.
- Force, F, is a result of mass, m, multiplied by acceleration, a ($F = ma$)

$$F = (10{,}000 \text{ kg})(9.8 \text{ m/s}^2) = 1 \times 10^5 \text{ kg-m/s}^2 = 1 \times 10^5 \text{ N}$$

- Newton, a unit of force, is equal to 1 kg-m/s^2

$$P = \frac{F}{A} = \frac{1 \times 10^5 \text{ N}}{1 \text{ m}^2} = 1 \times 10^5 \text{ N/m}^2 = 1 \times 10^5 \text{ Pa} = 1 \times 10^2 \text{ kPa}$$

UNITS OF PRESSURE

- There are many units of pressure. The SI unit for pressure is the pascal (Pa) which is 1 N/m^2.
- Because the pascal unit is small (e.g., atmospheric pressure is ~100 000 Pa = 100 kPa) the unit of bar is sometimes used.
  - Atmosphere (atm), torr, psi, kg/m^3 are additional units.

101325 Pa = 101.325 kPa = 1.01325 bar

MORE UNITS OF PRESSURE

- Scientists in the 17th century calculated the pressure exerted by the atmosphere using a 760 mm mercury filled tube. The mercury would be driven further into the tube by atmospheric pressure. Less mercury would indicate lower pressure.
- This lead to further units of pressure, 760 mm Hg and eventually torr. 1 torr = 1 mm Hg.

1 atm = 760. mm Hg = 760. torr = 1.01325 Pa = …
… 101.325 kPa = 1.01325 bar

	1 Bar	1 atm	1 Pa	1 kPa	1 Torr
Bar	1.000 Bar	1.01325 Bar	0.000010 Bar	0.0100 Bar	0.00133322 Bar
Atmosphere (atm)	0.987 atm	1.000 atm	0.00000987 atm	0.00986923 atm	0.0013158 atm
Pascal (Pa)	100,000 Pa	101325 Pa	1 Pa	1000 Pa	133.322 Pa
Kilopascal (kPa)	100 kPa	101.325 kPa	0.001 kPa	1.00 kPa	0.1333422
torr	750.062 torr	760 torr	0.00751 torr	7.50062 torr	1.00 torr

EXAMPLE: CONVERSIONS

a. Convert 0.357 atm to torr. **271 torr.**
b. Convert 6.6 x 10^{-2} torr to atmospheres **8.7x10^{-5} atm**
c. Convert 147.2 kPa to torr. **1104 torr**

QUESTIONS

 # 10.13, 10.21, 10.22

TOPIC ONE & TWO PRACTICE PROBLEMS

THE GAS LAWS

TOPIC THREE

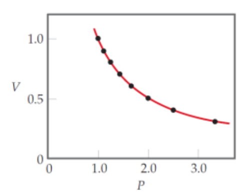
THE GAS LAWS

- Gases are defined by four variables which affect their physical condition or state.
  - **Temperature**
  - **Pressure**
  - **Volume**
  - **Number of Moles (amount of gas)**

PRESSURE-VOLUME

- Robert **Boyle** was one of the first chemists to investigate the relationship between **pressure and volume**.
- Using a J shaped tube filled with a gas he added mercury and *observed* that as the pressured increased the volume decreased.
- He measured that if the pressure was doubled, the volume of gas was reduced to half.

BOYLE'S LAW



- This graph can be read as 'As volume decreased, the pressure that a gas exerts increases'.
  - This assumes the temperature and amount of gas is held constant!

BOYLE'S LAW

- Boyle's Law states *the volume of a fixed quantity of gas maintained a constant temperature is inversely proportional to the pressure.*
- Inverse properties are those in which one gets smaller as the other increases.
  - E.g., as you move farther from the sun, the temperature of a planet decreases.

$$V = \text{constant} \times \frac{1}{P} \quad \text{or} \quad PV = \text{constant}$$

- The value of the constant depends on the temperate and amount of gas present.

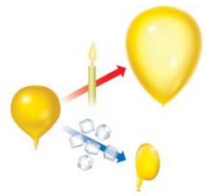
EXAMPLE: PRESSURE-VOLUME

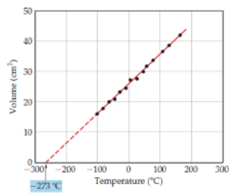
Predict, then explain, what happens to the pressure when you double the volume of a container maintaining a constant pressure and temperature.

The pressure would decrease by half.
Hint: Sub in 'fake' values for volume (v_1= 5, v_2= 10) and solve for P!

TEMPERATURE-PRESSURE

- The affect of temperature on pressure was investigated by Jacques Charles, and came to be understood through **Charles's Law**.
- He stated that *the* ***volume of a*** *fixed amount of* ***gas*** *maintained at constant pressure is* ***directly proportional to it's absolute temperature***.
  - In summary, gases will expand as they are heated from a point called absolute zero, a theoretical point at which gases will have zero volume. As temperature increases, volume increases. **This is a direct relationship**.
    - This is not achievable as all gases will liquefy or solidify before this point.





$$V = \text{constant} \times T \quad \text{or} \quad \frac{V}{T} = \text{constant}$$

QUANTITY-VOLUME

- Understood through the work of Gay-Lussac who observed that 1 volume of oxygen would react with 2 volumes of hydrogen to create 1 volume of water vapor.
- Furthered by Avogadro who proposed that equal volumes (constant temperature and pressure) of gases contain the same number of molecules.
- **Avogadro's Law states** ***the volume of a gas*** *maintained at constant pressure and temperature is* ***directly proportional to the number of moles of that gas.***
  - **In summary, doubling the number of moles of gas, doubles the volume the gas occupies.**

$$V = \text{constant} \times n$$

EXAMPLE: GAS LAWS

Consider a gas inside a container with a movable piston. Consider the changes if no gas is permitted to leak.

a. Heat the gas from 298 K to 360 K at constant pressure.

b. Reduce volume from 1.0 L to 0.5 L

c. Inject additional gas, keeping pressure and temperature constant.

Indicate how each change affects the average distance of the molecules, the pressure of the gas, and the number of moles of gas in the cylinder.

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ANSWERS

a. An increase in temperature, under constant pressure will cause the piston to move up therefore increases the volume. As the volume increases, the distance between the molecules increases. Pressure of the gas number of moles of gas remains unchanged.
b. Reduce volume (moves piston down), cause the gases to occupy less space, therefore a decrease in distance between molecules. As volume and pressure are inverse, as the volume decreases, pressure increases, while the number of moles remains unchanged.
c. An increase in number of moles decreases the distance between atoms (more atoms in the same amount of space), pressure increases.

QUESTIONS

#10.27, 10.28

TOPIC THREE PRACTICE PROBLEMS

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THE IDEAL-GAS EQUATION

TOPIC FOUR

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THE IDEAL-GAS EQUATION

- All three laws previously mentioned can be written as proportionality states as follows:

Boyle's law: $V \propto \frac{1}{P}$ (constant n, T)

Charles's law: $V \propto T$ (constant n, P)

Avogadro's law: $V \propto n$ (constant P, T)

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IDEAL-GAS LAW EQUATION

- We can combine these laws into one (left) , written as:

$$V \propto \frac{nT}{P} \qquad PV = nRT$$

- This states that the volume of a gas is proportional to the number of moles, temperature and pressure under specific conditions.
- When a constant of proportionality is added, R, we can rewrite this equation as *the* ***ideal gas equation*** (right)

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IDEAL-GAS EQUATION

$$PV = nRT$$

P – Pressure – units can vary. Commonly atm, bar or kPa.

V – Volume – unit is usually litres.

n = number of moles – measured in mol.

R – gas constant – varies depending on the other terms.

T – Temperature – always measured in kelvin.

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IDEAL GAS

- The ideal gas equation allows us to predict how a change in temperature, pressure, volume or number of moles will affect the system.
- This is based on an 'ideal' gas where:
  a. Molecules of the gas *do not* interact with each other.
  b. The volume of gas molecules themselves do not occupy any space in the container
- For the most part the error associated with these assumptions is small and therefore can be largely ignored.

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GAS CONSTANT

- The gas constant, R, is a quantity with several possible units as it is based on other measured quantities.
- Most commonly written as
  - 0.08206 L-atm/mol-K

TABLE 10.2 • Numerical Values of the Gas Constant *R* in Various Units

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m^3-Pa/mol-K*	8.314
L-torr/mol-K	62.36

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IDEAL-GAS EQUATION

- Determine the volume of 1.000 mol of gas at 0.0000 °C at 1.000 atm. **(22.41 L)**
- This is called standard temperature and pressure (STP), and is known as the molar volume of an ideal gas.

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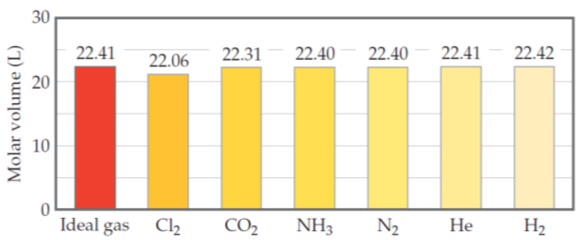
WHICH GAS IS THE CLOSEST TO IDEAL?



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EXAMPLE: IDEAL GAS EQUATION

- Calcium carbonate, in the main compound in limestone decomposes into calcium oxide and carbon dioxide. Sample of $CaCO_3$ is decomposed and the carbon dioxide is collected in a 250-mL fask. After decomposition is complete, the gas has a pressure of 1.3 atm at a temperature 31 C. How many moles of CO_2 gas was produced?

0.013 mol CO_2

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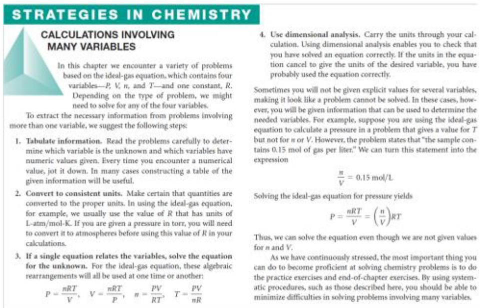
EXAMPLE: IDEAL-GAS EQUATION

Tennis balls are usually filled with air or nitrogen gas to a pressure of atmospheric pressure to increase their bounce. If a tennis ball has a volume of 144 cm^3 (144 mL) and contains 0.33 g of N_2 gas, what is the pressure inside the ball at 24 °C?

2.0 atm



STRATEGIES IN CHEMISTRY

CALCULATIONS INVOLVING MANY VARIABLES

In this chapter we encounter a variety of problems based on the ideal-gas equation, which contains four variables—P, V, n, and T—and one constant, R. Depending on the type of problem, we might need to solve for any of the four variables.

To extract the necessary information from problems involving more than one variable, we suggest the following steps:

1. **Tabulate information.** Read the problems carefully to determine which variable is the unknown and which variables have numeric values given. Every time you encounter a numerical value, jot it down. In many cases constructing a table of the given information will be useful.
2. **Convert to consistent units.** Make certain that quantities are converted to the proper units. In using the ideal-gas equation, for example, we usually use the value of R that has units of L-atm/mol-K. If you are given a pressure in torr, you will need to convert it to atmospheres before using this value of R in your calculations.
3. **If a single equation relates the variables, solve the equation for the unknown.** For the ideal-gas equation, these algebraic rearrangements will all be used at one time or another:

$$P = \frac{nRT}{V}, \quad V = \frac{nRT}{P}, \quad n = \frac{PV}{RT}, \quad T = \frac{PV}{nR}$$

4. **Use dimensional analysis.** Carry the units through your calculation. Using dimensional analysis enables you to check that you have solved an equation correctly. If the units in the equation cancel to give the units of the desired variable, you have probably used the equation correctly.

Sometimes you will not be given explicit values for several variables, making it look like a problem cannot be solved. In these cases, however, you will be given information that can be used to determine the needed variables. For example, suppose you are using the ideal-gas equation to calculate a pressure in a problem that gives a value for T but not for n or V. However, the problem states that "the sample contains 0.15 mol of gas per liter." We can turn this statement into the expression

$$\frac{n}{V} = 0.15 \text{ mol/L}$$

Solving the ideal-gas equation for pressure yields

$$P = \frac{nRT}{V} = \left(\frac{n}{V}\right)RT$$

Thus, we can solve the equation even though we are not given values for n and V.

As we have continuously stressed, the most important thing you can do to become proficient at solving chemistry problems is to do the practice exercises and end-of-chapter exercises. By using systematic procedures, such as those described here, you should be able to minimize difficulties in solving problems involving many variables.

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IDEAL-GAS & GAS LAWS

- If the number of moles, n, and temperature is held constant then a change in volume or pressure can be predicted using Boyle's Law (pressure and volume are inverse).
- This change can be calculated using the ideal-gas equation but because nRT will always yield a constant it can be written as:

$$P_1V_1 = P_2V_2$$

- This formula shows that if pressure decreases ($P_1 > P_2$), then the volume of the container has expanded ($V_2 > V_1$)

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IDEAL-GAS & GAS LAWS

- This derivation works for any of the properties of gases as long as the other two are held constant.
  - Possible derivations include:

$$P_1V_1 = P_2V_2 \qquad \frac{P_1}{n_1} = \frac{P_2}{n_2} \qquad \frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{V_1}{n_1} = \frac{V_2}{n_2} \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

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EXAMPLE: IDEAL-GAS & BOYLE'S LAW

The gas pressure in an aerosol can is 1.5 atm at 25 ° C. Assuming the gas obeys the ideal-gas equation, what is the pressure when the can is heated to 450 °C.

3.6 atm

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EXAMPLE: GAS LAW & BOYLE'S LAW

The pressure in a natural-gas tank is maintained at 2.20 atm. On a day when the temperature is -15 °C, the volume of the gas is 3.25 x10^3 m^3. What is the volume of the same quantity of gas on a day when the temperature is 31°C?

3.83 x 10^3 m^3

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COMBINED GAS LAW

- Similarly, when the number of moles is held constant, changes with pressure, temperature or volume can be predicted using a simplified formula.
- This formula is called the combined gas law:

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

EXAMPLE: COMBINED GAS LAWS

An inflated balloon has a volume of 6.0 L at sea level (1.0 atm) and is allowed to ascend until the pressure is 0.45 atm. During the ascent, the temperature of the gas falls from 22 ° C to -21 °C. Calculate the final volume of the balloon at it's final altitude.

11 L

EXAMPLE: COMBINED GAS LAWS

A 0.50-mol sample of oxygen gas is confined at 0 °C and 1.0 atm in a cylinder with a movable piston. The piston compresses the gas so that the final volume is half of the initial volume and the final pressure is 2.2 atm. What is the final temperature of the gas in degrees Celsius?

27 °c

QUESTIONS

#10.32, 10.35, 10.36, 10.37, 10.41

TOPIC FOUR PRACTICE PROBLEMS

GAS MIXTURES AND PARTIAL PRESSURES

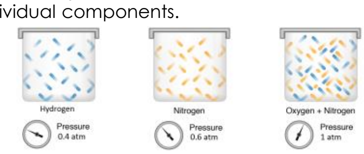
TOPIC FIVE

PARTIAL PRESSURES

- John Dalton (important to atomic theory) worked with gases being mixed.
- He discovered that: ***the total pressure of a mixture of gases is equal to the sum of the pressures that each would exert if it were present alone.***
- The pressure that a gas exerts in a mixture is called the **partial pressure**, sometimes known as **Dalton's Law of Partial Pressures**

PARTIAL PRESSURE

- If the pressure of a container (P_t) is due to a mixture of three gases (P_1, P_2, P_3) then the total pressure is equal to the sum of the individual components.



$$P_t = P_1 + P_2 + P_3$$

PARTIAL PRESSURES

- This concept implies that each gas is unaffected by the others and the pressure of each gas can be determined using the ideal-gas equation.

$$P_1 = n_1\left(\frac{RT}{V}\right); \quad P_2 = n_2\left(\frac{RT}{V}\right); \quad P_3 = n_3\left(\frac{RT}{V}\right); \quad \text{and so forth}$$

- Each gas is at the same temperature, pressure and volume which means it can be rewritten to determine the total pressure after the total number of moles is determined.

$$P_t = (n_1 + n_2 + n_3 + \cdots)\left(\frac{RT}{V}\right) = n_t\left(\frac{RT}{V}\right)$$

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EXAMPLE: PARTIAL PRESSURE

A mixture of 6.00 g oxygen gas and 9.00 grams of methane is placed into a15.0-L vessel at 0°C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

0.281 atm, 0.841 atm & 1.122 atm

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EXAMPLE: PARTIAL PRESSURE

What is the total pressure exerted by a mixture of 2.00 g of H_2(g) and 8.00 g N_2(g) at 273 K in a 10.0-L vessel?

2.86 atm

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PARTIAL PRESSURE & MOLE FRACTIONS

- Because each gas behaves nearly as an ideal gas, the amount of pressure exerted depends only on the number of moles present, not the type of compound.
- A mole fraction is a dimensionless number that indicates a ratio of the number of moles of a gas over the total of moles in the mixture given by the formula:

$$P_1 = \left(\frac{n_1}{n_t}\right)P_t = X_1 P_t$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

AP Format

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EXAMPLE: PARTIAL PRESSURE

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed on 1.5 mol percent CO_2, 18.0 mol percent O_2 and 80.5 mol percent Ar. (a) calculate the partial pressure of O_2 in the mixture if the total pressure is to be 745 torr. (b) If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O_2 are needed?

(a) 134 torr (b) 0.879 mol O_2

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EXAMPLE: PARTIAL PRESSURE

From data gathered by ***Voyager 1***, scientists have estimated the composition of the atmosphere of Titan, Saturn's largest moon. The pressure on the surface of Titan is 1220 torr. The atmosphere consists of 82 mol percent N_2, 12 mol percent Ar and 6.0 mol percent CH_4. Calculate the partial pressure of each gas.

1.0×10^3 torr N_2, 1.5×10^2 torr Ar & 73 torr CH_4

QUESTIONS

#10.63, 10.65, 10.66

TOPIC FIVE PRACTICE PROBLEMS

KINETIC MOLECULAR THEORY

TOPIC SIX

KINETIC MOLECULAR THEORY

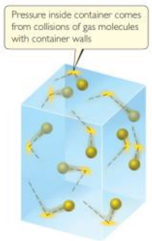
- The ideal gas law is able predict how a change to the properties of gases will affect the system but it does not ***explain*** how this occurs.
- The model used today, known as the **kinetic-molecular theory** of gases, helps us to picture what happens to gas particles when conditions change.

KMT UNDERSTANDINGS

1. Gases consists of large number of molecules always in random motion.
2. The molecules themselves are negligible relative to the total volume (they take up almost zero space).
3. The molecules do not attract or repel one another (no interactions)
4. Energy can be transferred between molecules through collisions but at a constant temperature the average kinetic energy remains the same.
5. The average kinetic energy is proportional to the temperature. At any given temperature, all molecules have the same average kinetic energy.

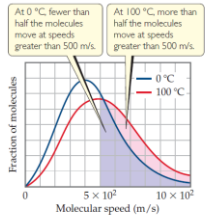
PRESSURE AND KMT

- The pressure of a gas is caused by the collisions of molecules with the sides of the container.
- The pressure is determined by the frequency and force of the collisions.
- A rise in temperature causes a rise in pressure as molecules have greater kinetic energy and strike the container more frequency.



MOLECULAR SPEED

- The speed at which molecules move is a function of their kinetic energy. While this energy can be transferred, the total kinetic energy remains stable (at constant temperature).
- Because collision will result in large variance of kinetic energy (one moves away with greater, while the other is much slower) the molecular speeds of a sample can have great variance.
  - The higher the temperature, the greater the average kinetic energy.



APPLICATION OF KMT

1. **An increase in volume at a constant temperature causes pressure to decrease.** This decrease in pressure is due to the greater distance the molecules must travel before the strike the container. Their kinetic energy remains constant under constant temperature.
2. **A temperature increase at a constant volume causes pressure to increase.** As molecules move faster, and there is no change in volume, there are more collisions per unit of time with the walls of the container. Secondly, the average kinetic energy increases which means each collision with the walls of the container is more forceful. Both of these aspect cause the pressure to increase.

QUESTIONS

#10.77, 10.79

TOPIC SIX PRACTICE PROBLEMS

IDEAL GASES VS. REAL GASES

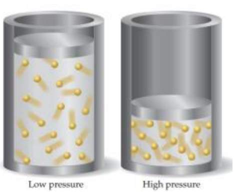
TOPIC SEVEN

REAL GASES

- Real gases differ from ideal gases. While the difference under normal conditions (small pressures – under 10 atm) and temperatures much greater then 0K.
- Under these conditions the ideal-gas equation works well, but as the conditions become more extreme (extremely low temperature/high pressure) then the error increases.

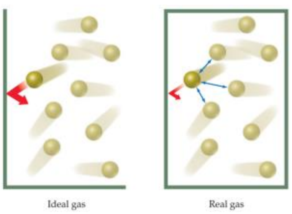
IDEAL GAS BEHAVIOR

- Ideal gases are based on the assumption that the particles themselves do not contribute to the volume of the gas, and they do not attract either other.
- At **extremely high pressures** (low volume) the molecules themselves begin to contribute to the space and cannot be ignored as they are in low pressure systems.



IDEAL GAS BEHAVIOR

- Secondly, ideal gas molecules do not have attractive/repulsive forces.
- In real gases, under high pressure conditions, the molecules are much closer together and **attract slightly** to one another which lowers the force of the collision when it collides with the container wall.



QUESTION
#10.93(A&B)

TOPIC SEVEN PRACTICE PROBLEMS