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Where Minds Pollinate

Class 12
Physics

Chapter 16

Exercise Solution

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Short Questions

1) Mention the different ways of increasing the number of molecular collisions per unit time in a gas.

There are three primary ways to increase the number of moleculer collisions per unit time in a gas:

1. Increase temperature: Raising the temperature increase the average kinetic energy and thus the average speed of gas molecules. Faster-moving molecules collide more frequently with each other and the container walls.
2. Increase Pressure (Decrease Volume): Compresing the gas into smaller volume means the molecules are packed more closely together. This reduces the mean free path, leading to more frequent collisions.
3. Increase the number of molecules: Adding more gas molecules to the same container means there are simply more particles to collide, leading to a higher collision frequency.

2) By reducing the volume of gas at a constant temperature, the pressure of the gas increases. Explain it on the basis of kinetic theory.

When the volume of a gas is reduced at a constant temperature, the molecules are confined to a smaller space. Since the temperature is constant, the average kinetic energy and thus the average speed of molecules remain unchanged. However, because the molecules are now in a smaller volume, they will collide with the container walls more frequently per unit time. Each collision transfers momentum to the wall, and an increased frequency of these momentum transfers results in a greater average force exerted on the walls per unit area, which is observed as an increase in pressure.

3) What do you mean by the root mean square speed of the molecules of gas? Is the root mean square speed the same as the average speed?

The root mean square (rms) speed is the square root of the averge of the squares of the molecular speeds in a gas. It gives a measure of the average kinetic energy of gas molecules.

Mathematically: $C_{rms} = \sqrt{3KT/m} = \sqrt{3RT/M}$

Root mean square speed of the molecule of a gas is not same as the average speed. Rms speed is slightly higher than average speed. Both describe molecular motion but are calculated differently.

4) Why is the temperature below absolute zero not possible?

Temperature below absolute zero is not possible because at absolute zero (0K), the molecular motion stops completly, and the gas has minimum internal energy. Temperature is a measure of this kinetic energy. If a system were to go below absolute zero, its particles would theoretically have less than zero kinetic energy, which is physically impossible. You cannot have "less" motion than no motion.

5) Estimate the average energy of a helium atom at the temperature on the surface of the sun (6000K)

Data:

$KE_{av} = ?$

$T = 6000K$

Solution:

The average energy of a gas atom is

$KE_{av} = 3/2\, KT$

$KE_{av} = 3/2\,(1.38 \times 10^{-23})(6000)$

$K.E_{av} = 1.2 \times 10^{-19} J$

6. Show that the ratio of the root mean square speeds of molecules of two different gases at a centaine temperature is equal to the square root of the inverse ratio of their masses.

The root mean square of gas molecules is given by:

$C_{rms} = \sqrt{3KT/m}$

For two gases with molecular mass m_1 and m_2

$$\frac{C_{rms1}}{C_{rms2}} = \frac{\sqrt{3KT/m_1}}{\sqrt{3KT/m_2}}$$

$$\frac{C_{rms1}}{C_{rms2}} = \sqrt{\frac{m_2}{m_1}}$$

This shows that the RMS speed ratio is equal to the square root of the inverse ratio of their masses.

7) Differentiate between the formations of white dwarfs and neutron stars.

	White Dwarf	Neutron Stars
Formation	From low/mid mass stars	From massive stars
Core Composition	Degenrate electron gas	Degenrate neutrons
Density	$\sim 10^6 g/cm^3$	$\sim 10^{14} g/cm^3$
Size	Earth-sized (~6,000 km diameter)	Tiny (~10-20 km diameter)
Rotation	Slow (hours/days)	Extremely fast (ms to s)
Magnetic field	Moderate ($\angle$1 million Gauss)	Extreme (10^{12} – 10^{15} Gauss)
Energy Emission	Fades Slowly	Emits X-rays/pulsar beams.
Example	Sirius B	Crab Pulsar

8) Why do the gases at low temperatures and high pressure show large deviations from ideal behaviour?

Gases deviate significantly from ideal behaviour at low temperatures and high pressures due to:

1. Intermolecular Forces: At low temps, molecules move slower, allowing attracting forces (Van Der Wal) to dominate, reducing pressure.
2. Finite Molecular Volume: At high pressures, molecules are forced closer, making their actual volume significant compared to container size.

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9) What distinguishes degenerate matter from regular matter?
Degenrate matter differs from regular matter in these key ways:
1. Pressure Source:
• Regular Matter, pressure arises from thermal motion (kinetic energy).
• Degenrate Matter, pressure comes from quantum mechanical effect, independent of temperature.
2. Density: Degenerate matter is ultra-dense.
3. Temperature Sensitivity:
• Regular matter expands when heated.
• Degenerate matter's pressure not changes with temperature.
Example: White dwarfs (electron-degenrate), Neutron Stars (neutron-degenrate).
10) Show that the temperature of ideal gas is directly propotional to the average translational kinetic energy of gas molecules.
The equation of the pressure of gas is:
$P = 1/3\ mN/V \langle V^2 \rangle$
$PV = mN/3 \langle V^2 \rangle$
As we know that,
$PV = NKT$
So,
$mN/3 \langle V^2 \rangle = NKT$
$1/3\ m \langle V^2 \rangle = KT$
Multiplying both sides by 3/2
$3/2\ (1/3\ m \langle V^2 \rangle) = 3/2\ KT$
$1/2\ m \langle V^2 \rangle = 3/2\ KT$
$K.E_{av} = (\text{constant})T$
$K.E_{av} \propto T$
11) What happens to the electrical resistance of a superconductor when it is cooled below its critical temperature?
When cooled below its critical temperature, the electrical resistance of a superconductor drops abruptly to zero, allowing it to conduct electricity with no energy loss. This phenomenon, known as superconductivity, results in the loss of all resistance, meaning current can flow indefinity with dissipation.
12) What are some potential applications of superconductors in transportation?
Superconductors can be used in magnetic leviation (Maglev) trains. They reduce friction and allows train to float and move smoothly. Superconductors are also useful in high-speed electric motors and hyperloop systems. They help in efficient power transfer and fast braking systems.
13) Provide an example of a high-temperature superconductor.
Yttrium Barium Copper Oxide (YBCO). Its chemical formula is $YBa_2Cu_3O_7$. It becomes superconducting above 90K, which is above boiling point of liquid nitrogen. It is widely used in reasearch and technology.

Conceptual Questions:

1) What do you mean by elastic collision of gas molecules?
An elastic collision of gas molecules is one where the total kinetic energy of the colliding molecules remains constant both before and after collision, with no energy lost to other forms like heat or sound.
2) Why pressure is equally distributed on all faces ($P_x = P_y = P_z = P$)?
Pressure in a gas is equally distributed ($P_x = P_y = P_z = P$) because gas molecules are in constant, random, and isotropic (equal in all directions) motion, leading to uniform force and pressure on all surfaces of a container. The cumulative effect of these countless, random molecular collisions results in an average, constant pressure that acts perpendicular to the container walls in all directions at any given point with the gas.
3) Why do we consider mean square velocity?
We consider mean square velocity because it provides a relevent measure of the total kinetic energy of a system of particles, like gas molecules, which is often more useful than the simple average velocity. The average velocity of particles moving in random directions is zero, but the mean square velocity is always non-zero and directly relates to the kinetic energy, temperature, and pressure of the system.
4) State a relation between pressure and kinetic energy (K.E)?
As we know that
$P = 1/3\ Nm/V \langle V^2 \rangle$ — (i)
And, $K.E = 1/2\ mv^2$
Multiplying and dividing by 2 in equation (i)
$P = 2/3\ N/V \langle 1/2\ mv^2 \rangle$
$P = 2/3\ N_0 \langle K.E \rangle$ $\quad \because N/V = N_0$
$P = \text{Constant} \langle K.E \rangle$
$P \propto \langle K.E \rangle$
5) Derive relation between pressure and density?
As we know that,
$P = 1/3\ Nm/V \langle V^2 \rangle$
Total mass = Nm
Total volume = V
Density = ρ = mass / volume = Nm/V
So,
$P = 1/3\ \rho \langle V^2 \rangle$
6) Why is $\langle V_x^2 \rangle = \langle V_y^2 \rangle = \langle V_z^2 \rangle$?
The equation $\langle V_x^2 \rangle = \langle V_y^2 \rangle = \langle V_z^2 \rangle$ implies that the overge square of the velocity components along the x, y and z axes are equal, which arises from the principle of equipartioning of energy in thermodynamics, where in a system with

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no preferred direction, like in a gas, the total kinetic energy is equally distributed among the three degrees of freedom (x, y and z) directions. This results in the average kinetic energy for each direction being equal.

7) How temperature is related to the average kinetic energy of molecules?

As we know that,

$P = 1/3\ Nm/V \langle v^2 \rangle$

$PV = 1/3\ Nm \langle v^2 \rangle$

Multiplying and dividing by 2

$PV = 2/3 N \langle 1/2\ mv^2 \rangle$

$PV = 2/3 N \langle K.E \rangle$ — (i)

General gas equation

$PV = nRT$ $\because n = N/N_A$

$PV = \frac{NRT}{N_A}$ $\because K = R/N_A$

$PV = NKT$ — (ii)

Comparing equation (i) and equation (ii)

$NKT = 2/3 N \langle K.E \rangle$

$KT = 2/3 \langle K.E \rangle$

$T = 2/3K \langle K.E \rangle$

$T = \text{Constant} \langle K.E \rangle$

$T \propto \langle K.E \rangle$

8) If we increase number of particles then what will be the affect on pressure?

Increasing the number of gas particles in a container will increase the pressure. This is because more particles lead to higher frequency of collisions with the container walls, and gas pressure is essentially the force of these collisions per unit area.

9) RMS velocity of a gas increases with temperature. Justify?

The root-mean-square (rms) velocity of gas molecules increases with temperature because temperature is a direct measure of the average kinetic energy of those molecules. Higher temperatures mean particles have more kinetic energy, and since kinetic energy is propotional to the square of velocity ($1/2\ mv^2$) they must move faster on average to achieve this higher energy. This is reflected in the equation for rms velocity ($\sqrt{3RT/M}$), which shows a direct propotionality between V_{rms} and the square root of the absolute temperature.

10) What is physical significance of root mean square value of velocity of gas?

The root mean square velocity of a gas ~~molecule~~ is a measure of the efective average speed of its molecules, representing the speed of hypothetical molecule would have if it possessed the average kinetic energy of the entire gas sample. It's physical significance lies in providing a non-zero value for the molecular motion that directly relates to the gas's temperature and kinetic energy, which are macroscopic properties crucial in understanding and predicting the gas's behaviour.

11) Why we take rms value of velocity of molecules of gas?

We use the root mean square (rms) velocity for gas molecules because a simple average velocity would be zero, a gas particles move randomly in all directions. The rms velocity is a measure of the average speed that captures the kinetic energy of the gas, is always non-zero, and is essential for calculating diffusion rates, effusion rates and for understanding the gas's pressure and temperature in kinetic theory.

12) Why do real gas deviate from ideal gas behavior?

Real gases deviate from ideal behaviour because their molecules have finite volume and exert intermolecular forces, which are not negligible at high pressures and low temperatures. Unlike ideal gases, which are assumed to have zero volume and no intermolecular forces, real gas molecules occupy space and interact with each other, causing deviations from the ideal gas law.

13) Under what conditions real gas behaves like an ideal gas?

A real gas behaves like an ideal gas under conditions of low pressure and high temperature. At low pressure, gas molecules are far apart, minimizing intermolecular forces and their individual volumes becoming insignificant compared to the total space. High temperatures provide the molecules with high kinetic energy, making them move quickly and further reducing the impact of these intermolecular forces.

14) Why does pressure of real gas is less than predicted by ideal gas?

The pressure of a real gas is less than predicted by the ideal gas law because of intermolecular attractive forces that pull gas particles inward, reducing the force of their collisions with the container walls, and to a lesser extent, the actual volume of gas molecules which occupies some of the total volume.

15) What do you mean by microsocpic and macroscopic properties of gas?

Microscopic properties of gases describe the behaviours of individual atoms and molecules, such as their speed, momentum and kinetic energy, while macroscopic properties describe the observable characteristics of the gas as a whole, such as its pressure, volume and temperature.

16) What are the limitations to use $PV = nRT$ equation?

The $PV = nRT$ equation has limitations because it is based on idealized behaviour of gasses, which isn't real for several reasons. The main limitations are that the equation fails to account for intermolecular forces and the finite volume of gas molecules.

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Numericals

1) The mass of a helium atom is 6.64×10^{-27} kg. Calculate the root mean square speed of helium atom in a gas at a temperature of 15°C.

Data:

$m = 6.64 \times 10^{-27}$

$T = 15°C = 288K$

$C_{rms} = ?$

Solution

$C_{rms} = \sqrt{\frac{3KT}{m}}$

$C_{rms} = \sqrt{\frac{3(1.38 \times 10^{-23})(288)}{6.64 \times 10^{-27}}}$

$C_{rms} = \sqrt{179.6 \times 10^4}$

$C_{rms} = 1.34 \times 10^3 \, ms^{-1}$

2) At which temperature will the root mean square velocity of the oxygen molecules become equal to the escape velocity of the earth ($11.2 \, kms^{-1}$)? (Mass of one molecule of oxygen is 5.3×10^{-26} kg) (Boltzman Constant k is $1.38 \times 10^{-23} \, JK^{-1}$)

Data:

$T = ?$

$C_{rms} = V_e = 11.2 \, kms^{-1} = 11200 \, ms^{-1}$

$m = 5.3 \times 10^{-26}$ kg

$k = 1.38 \times 10^{-23} \, JK^{-1}$

Solution:

$C_{rms} = \sqrt{\frac{3kT}{m}}$

$C_{rms}^2 = \frac{3kT}{m}$

$T = \frac{C_{rms}^2 m}{3k}$

$T = \frac{(11200)^2(5.3 \times 10^{-26})}{3(1.38 \times 10^{-23})}$

$T = 1.6 \times 10^5 K$

3) The mass of molecule of a gas is 6.4×10^{-27} kg. Calculate the root mean square speed of gas molecule and kinetic energy per molecule at temperature 400K.

Data:

$m = 6.4 \times 10^{-27}$ kg

$C_{rms} = ?$

$K.E = ?$

$T = 400K$

Solution:

$C_{rms} = \sqrt{\frac{3KT}{m}}$

$C_{rms} = \sqrt{\frac{3(1.38 \times 10^{-23})(400)}{6.4 \times 10^{-27}}}$

$C_{rms} = \sqrt{258.75 \times 10^4}$

$C_{rms} = 1.6 \times 10^3 \, ms^{-1}$

$K.E = 3/2 \, KT$

$K.E_{av} = 3/2 \, (1.38 \times 10^{-23})(400)$

$K.E_{av} = 8.28 \times 10^{-21} J$

4) At which temperature the rms speed of gas molecules becomes double than its speed at 27°C., (pressure of gas is kept constant).

Data:

$T_2 = ?$

$C_{rms2} = 2C_{rms1}$

$T_1 = 27°C = 27 + 273 = 300K$

Solution:

The rms speed is propotional to the square root of temperature.

$C_{rms} \propto \sqrt{T}$

$\frac{C_{rms2}}{C_{rms1}} = \frac{\sqrt{T_2}}{\sqrt{T_1}}$

$\frac{2C_{rms1}}{C_{rms1}} = \frac{\sqrt{T_2}}{\sqrt{T_1}}$

$2 = \frac{\sqrt{T_2}}{\sqrt{300}}$

$2(17.32) = \sqrt{T_2}$

$\sqrt{T_2} = 34.64$

$T_2 = 1200K = 927°C$

5) Determine the root mean square speed of argon atoms at temperature 40°C. The molar mass of argon is $39.95 \, gmol^{-1}$

Data:

$C_{rms} = ?$

$T = 40°C = 313K$

$M = 39.95 \, gmol^{-1} = 0.03995 \, kgmol^{-1}$

Solution

$C_{rms} = \sqrt{\frac{3RT}{M}}$

$C_{rms} = \sqrt{\frac{3(8.314)(313)}{0.03995}}$

$C_{rms} = 442 \, ms^{-1}$

6) Calculate the number of gas molecules in a cubic meter of gas at standard temperature and pressure (STP).

Data:

$m = ?$

$V = 1m^3 = 1000L$

$T = 273K$

$P = 1.01 \times 10^5 Pa$

Solution:

Use the Ideal Gas Law to find moles of gas

$PV = nRT$

$n = PV/RT$

$n = \frac{(1.01 \times 10^5)(1)}{(8.314)(273)}$

$n = 44.5 \, mol$

Convert Moles to Number of molecules.

$m = nN_A$

$m = (44.5)(6.022 \times 10^{23})$

$m = 2.68 \times 10^{25}$ molecules.

7) A vessel A contains hydrogen and another vessel B whose volume is twice of vessel A contains same mass of oxygen at the same temperature. Find (i) the ratio of the root mean square speeds of hydrogen and oxygen gases. (ii) the ratio of pressure of gases in vessels A

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and B. Molecular mass of hydrogen and oxygen are 2 and 32 respectively.

Data:

$V_A = V$

$V_B = 2V$

$\frac{C_{rmsH}}{C_{rmsO}} = ?$

$\frac{P_A}{P_B} = ?$

$M_H = 2\,g\,mol^{-1}$

$M_O = 32\,g\,mol^{-1}$

Solution:

The root mean square of gas is given by

$$C_{rms} = \sqrt{\frac{3RT}{M}}$$

At constant temperature

$$C_{rms} \propto \frac{1}{\sqrt{M}}$$

$$\frac{C_{rmsH}}{C_{rmsO}} = \frac{1/\sqrt{M_H}}{1/\sqrt{M_O}}$$

$$\frac{C_{rmsH}}{C_{rmsO}} = \frac{\sqrt{M_O}}{\sqrt{M_H}}$$

$$\frac{C_{rmsH}}{C_{rmsO}} = \frac{\sqrt{32}}{\sqrt{2}}$$

$$\frac{C_{rmsH}}{C_{rmsO}} = \frac{4}{1}$$

The ideal equation is given by

$$PV = nRT$$

$$P = nRT/V$$

$$\frac{P_A}{P_B} = \frac{n_A RT/V_A}{n_B RT/V_B}$$

$$\frac{P_A}{P_B} = \frac{n_A/V_A}{n_B/V_B}$$

$$\frac{P_A}{P_B} = \frac{n_A V_B}{V_A n_B}$$

Putting the values of V_A and V_B

$$\frac{P_A}{P_B} = \frac{n_A 2V}{V n_B}$$

$$\frac{P_A}{P_B} = \frac{2n_A}{n_B}$$

$$\frac{P_A}{P_B} = \frac{2m/M_H}{m/M_O}$$

$$\frac{P_A}{P_B} = \frac{2mM_O}{M_H m}$$

$$\frac{P_A}{P_B} = \frac{2M_O}{M_H}$$

$$\frac{P_A}{P_B} = \frac{2(32)}{2}$$

$$\frac{P_A}{P_B} = \frac{32}{1}$$

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