**The Kinetic Molecular Theory of Gases**

The laws that describe the behaviour of gases were well established long before anyone had developed a coherent model of the properties of gases. In this section, we introduce a theory that describes why gases behave the way they do.

**A Molecular Description**

The Kinetic Molecular Theory (KMT) of gases explains the laws that describe the behaviour of gases. It was developed during the mid-19th century by several physicists, including the Austrian Ludwig Boltzmann (1844–1906), the German Rudolf Clausius (1822–1888), and the Englishman James Clerk Maxwell (1831–1879). The kinetic molecular theory is based on the proposed properties of molecules of an ideal gas and provides a molecular explanation for observations that led to the development of the ideal gas law.

The kinetic molecular theory of gases is based on the following five postulates:

**1.** A gas is composed of a large number of molecules (atoms if a noble gas) that are in constant random motion.

**2.** Because the distance between gas molecules is much greater than the size of the molecules, the volume of the molecules themselves is negligible.

**3.** Intermolecular forces of attraction are so weak that they are also negligible.

**4.** Gas molecules collide with one another and with the walls of the container, but these collisions are perfectly elastic; that is, they do not change the kinetic energy of the molecules. Collision between the gas particles and the wall of a container cause pressure.

**5.** The average kinetic energy of the molecules of any gas depends on only the temperature. The higher the temperature the higher the average kinetic energy of the gas particles. The velocity of all particles increases at higher temperatures, and vice versa.

Although the molecules of real gases have nonzero volumes and exert both attractive and repulsive forces on one another, the kinetic molecular theory of gases can be used to explain the behaviour of gases under many real-life conditions – especially at low pressure and high temp. At low pressure and high temp the particles of real gases act very like ideal gas particles.

Postulates 1 and 4 state that gas molecules are in constant motion and collide frequently with the walls of their containers. The collision of molecules with their container walls results in a force exerted by the gas on the walls. Anything that increases the frequency with which the molecules strike the walls, or increases the momentum of the gas molecules (i.e., how hard they hit the walls), increases the pressure. Therefore - changing the concentration of the gas affects the frequency of collisions, while changing the temperature affects both frequency and momentum of the collisions.

Because volumes and intermolecular interactions are negligible, postulates 2 and 3 state that all gaseous particles behave identically, regardless of the chemical nature of their component molecules. This is the essence of the ideal gas law, which treats all gases as collections of particles that are identical in all respects except mass.

Postulate 2 also explains why it is relatively easy to compress a gas; you simply decrease the distance between the gas molecules.

Pressure versus Volume

Assuming a constant temperature, the average kinetic energy of the molecules of a gas and hence the average speed remains unchanged. Therefore if a given gas sample is allowed to occupy a larger volume, the density of the gas (number of particles per unit volume) decreases, and the average distance between the molecules increases. Hence the molecules must therefore collide with one another and with the walls of their containers less often, leading to a decrease in pressure. Conversely, compressing a gas forces the molecules closer together and increases the density, causing more collisions of the molecules with the container walls. When this increase of internal pressure balances the applied pressure used to compress the gas – the gas will no longer compress.

Volume versus Temperature

Raising the temperature of a gas increases the average kinetic energy and therefore the average speed of the gas molecules. Hence as the temperature increases, the molecules collide with the walls of their containers more frequently and with greater force. This increases the pressure, unless the volume increases to reduce the pressure, as we have just seen.

**Questions:**

1. Use your knowledge of the Kinetic Theory of Gases to explain why blowing air into a balloon cause the volume of the ball to increase. (Hint: part of this question is understanding which postulates of the Kinetic Theory are relevant to this question. Be really obvious that you know them by quoting them)
2. If the inflated balloon was tied off and placed in a freezer, it’s volume would decrease. Use your knowledge of the Kinetic Theory to explain why this would happen. Same hint as above.
3. If equal amounts of two different gases was placed in two identical balloons, would the two balloons have the same volume? Use your knowledge of the Kinetic Theory to explain your answer.