COMPRESSIBILITY FACTOR

The compressibility factor serves as a crucial determinant indicating the degree of deviation exhibited by real gases from the behavior predicted by the ideal gas model. Mathematically expressed as the ratio of the product of pressure (P) and volume (V) to the product of the number of moles (n), the ideal gas constant (R), and temperature (T), the compressibility factor is denoted by 'Z' and is defined as:

$$Z = \frac{PV}{nRT}$$

To assess the behavior of various gases, a graph is constructed, illustrating the relationship between the compressibility factor and pressure over a specified range. This graphical representation aids in comprehending and comparing how different gases deviate from ideal gas behavior under diverse pressure conditions. The compressibility factor becomes a valuable parameter for understanding real gas characteristics and predicting their behavior under varying pressures.

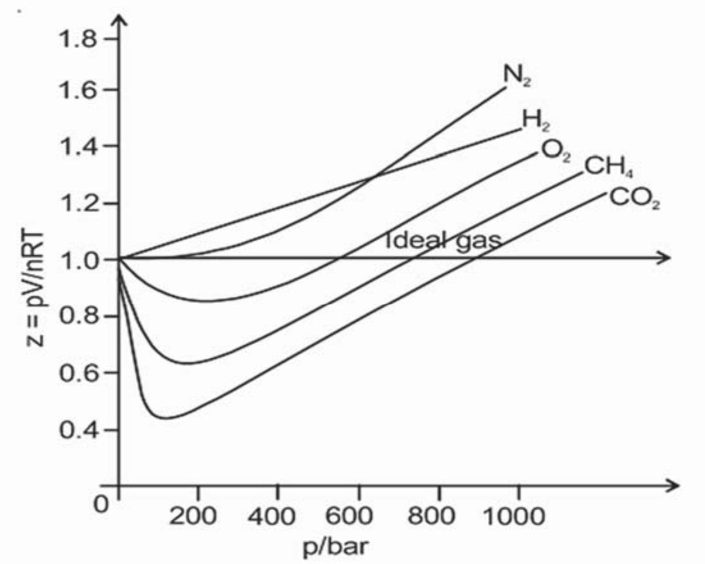

Fig.: Variation of compressibility factor for some gases

Conclusions made from the graph:

(i) **Z = 1:** The compressibility factor attains unity in the case of ideal gases across a wide spectrum of temperature and pressure conditions. This phenomenon is a consequence of the gas equation PV = nRT, where the equality of PV and nRT leads to a unity ratio.
Reason: In ideal gases, the gas equation is represented by PV = nRT. Since the product of PV is equivalent to nRT, the ratio of PV to nRT is invariably unity.

(ii) **Z ≈ 1:** The compressibility factor approaches unity under conditions of low pressure. This tendency arises due to the negligible interactions between molecules at low pressure, resulting in minimal impact on volume. Consequently, Z approximates 1.
Reason: At low pressure, the interactions among molecules are negligible, leading to an inconsequential effect on volume, and therefore Z tends to equal 1.

(iii) **Z > 1:** The compressibility factor surpasses unity, indicating positive deviation, particularly under high-pressure circumstances. The proximity of gas molecules intensifies at high pressure, prompting the onset of repulsive forces between them. While ideal gases exhibit a proportional decrease in volume with increasing pressure ($P \propto 1/V$), maintaining a constant product of PV, real gases deviate from this behavior. In real gases, the forces of repulsion hinder a proportional decrease in volume under high pressure, causing the product PV to increase with rising pressure ($PV > 1$). This concept is further elucidated through the following derivation:

$Z = PV_{real}/nRT$ ...(i)

Gas exhibiting ideal behavior is described by $V_{ideal} = nRT/P$ ...(ii)

By substituting the value of nRT/P from equation (ii) into equation (i),

we obtain $Z = V_{real} / V_{ideal}$. Thus, the compressibility factor (Z) is contingent upon the ratio of the actual volume of the gas to the volume of the ideal gas at the given temperature.

(iv) **$Z < 1$:** The compressibility factor falls below unity, indicative of negative deviation, when the pressure is in the intermediate range.

Reason: At intermediate pressure levels, gas molecules maintain a sufficient distance to mitigate repulsive forces. Instead, attractive forces come into play, leading to molecular attraction and closer proximity. Consequently, the volume experiences a more pronounced decrease than anticipated, and the reduction in volume is not proportionate to the increase in pressure. This results in a decrease in the product PV with rising pressure ($PV < 1$). Consequently, the ratio of PV/nRT becomes less than 1. Under such pressure conditions, gases exhibit greater compressibility owing to the prevalence of attractive forces. This underscores the conclusion that gas behavior approaches ideality when pressure is very low. Consequently, gases demonstrate ideal behavior when the occupied volume is extensive, allowing the volume of the gas molecules to be disregarded in comparison.

	Compressibility Factor Z	**Pressure**	**Compressibility**
(1)	$Z = 1$	All	Normal
(2)	$Z \approx 1$	Low	Normal
(3)	$Z > 1$	High	Difficult
(4)	$Z < 1$	Intermediate	Easy

It is noteworthy that pressure alone does not singularly determine the behavior of gases; temperature also plays a significant role.

Boyle Temperature or Boyle Point:

The Boyle Temperature, or Boyle Point, signifies the temperature at which real gases conform to ideal gas laws over a significant spectrum of pressure levels. The determination of Boyle Temperature is contingent upon the inherent properties and characteristics of the specific gas in question.

Effect of Temperature on the Compressibility Factor:

(i) Beyond the Boyle temperature: Real gases demonstrate positive deviation ($Z > 1$) from ideality. This phenomenon occurs as the temperature increases, causing the molecules to move farther apart. Consequently, the volume of the gas expands, leading to a reduction in the strength of the forces of attraction between the molecules.

(ii) Below the Boyle temperature: When operating below the Boyle temperature, the 'Z' value initially decreases and reaches a minimum with an increase in pressure. This decrease is attributed to the onset of attractive forces between the molecules. Subsequently, with a further increase in pressure, repulsive forces come into play, leading to a continuous increase in the 'Z' value.

Conclusions:

Gases exhibit ideal behavior under the following conditions:

(i) High temperature

(ii) Low pressure