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Real gas  $\Rightarrow$  Ideal gas & a Real gas.

\* Kinetic theory of gases  $\rightarrow$

- (1) There are no attractive or Repulsive forces b/w the gas molecule
- (2) There are no  $\Rightarrow$  volume of gas molecules,

Volume of Gas molecules is negligible as compared to the total volume of gas

# In Reality.

- a) If Gas volume is very small , molecules are close to each other,  $\bar{e}$  cloud of 1 molecule will repel the  $\bar{e}$  cloud of another. Repulsive forces are dominant.
- b) If Gas volume is large , then +ve charge nucleus of 1 molecule will attract the  $\bar{e}$  cloud of another which implies there will be attractive forces b/w the Gas molecules.
- c) If Gas volume is very large , then only we can neglect the volume of Gas molecules as compared to the total volume of Gas.
- d) If Gas has very small volume, then, we can not neglect the volume of gas molecule.

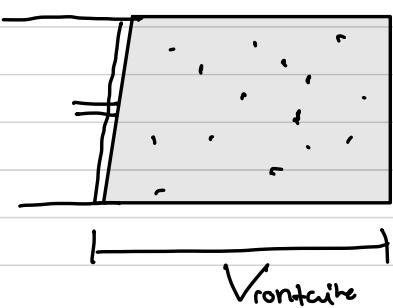
For Real gases we need to take into account the

- Attractive forces b/w gas molecules
- Repulsive forces b/w gas molecules
- Molecular volume of gas molecules

#

### 1. Ideal Volume

$\Rightarrow$  Free volume or net compressible volume is known as ideal volume.

# Ideal gas

$$V_{\text{ideal}} = V_{\text{container}}$$

# 2. Real Volume

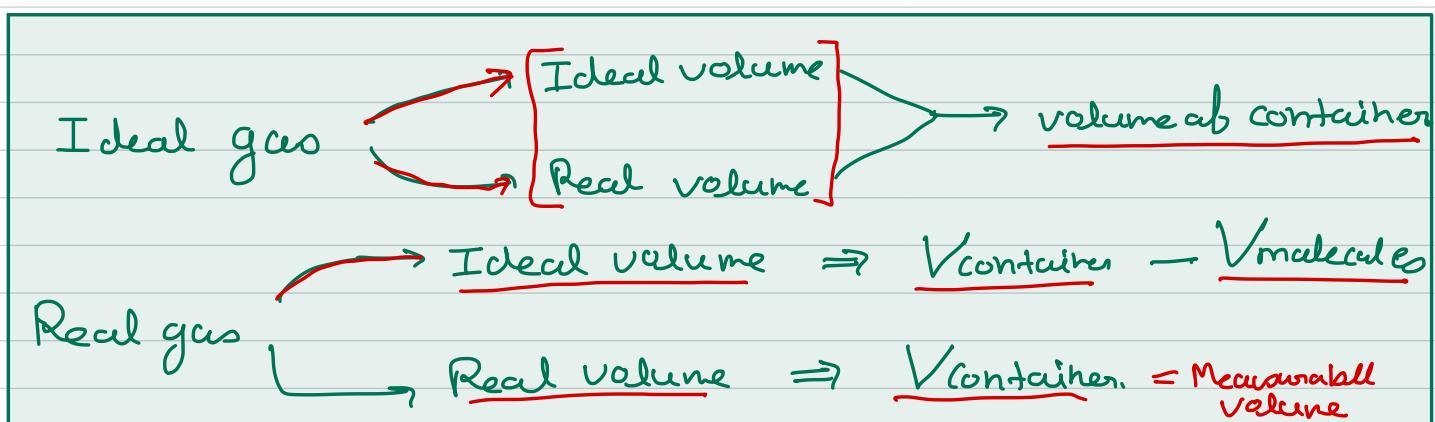
$\Rightarrow$  volume ab container

Measured volume

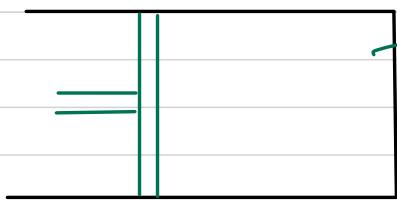
# What is the ideal volume of a Real gas

Compressible volume = Ideal volume

$$\Rightarrow V_{\text{container}} - V_{\text{molecules}}$$



(1)



Initial  
A(g)  
 $P_0, V_0$

$$P_f = 2P_0$$

$$V_f = \frac{V_0}{2}$$

we made the piston move inside till the final pressure became  $2P_0$

(2)



B(g)       $P_0, V_0$

$$P_f = 2P_0$$

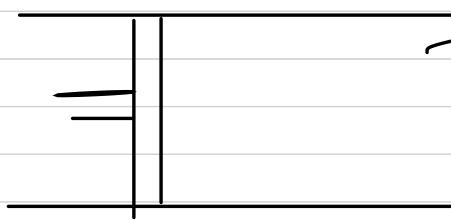
$$V_f > \left(\frac{V_0}{2}\right)$$

can i say that

A(g) is ideal gas

B(g) & C(g) are Real gas.

(3)



C(g)       $P_0, V_0$

$$P_f = 2P_0$$

$$V_f < \left(\frac{V_0}{2}\right)$$

B(g)  $(V_f) > \left(\frac{V_0}{2}\right)$

is less compressible as compared to ideal gas  $\Rightarrow$  In this case Repulsive forces must be dominant.

C(g)  $V_f < \left(\frac{V_0}{2}\right)$  at  $P_f = 2P_0$

is more compressible as compared to ideal gas  $\Rightarrow$  In this case attractive forces are dominant.

Compressibility factor  $\Rightarrow Z \Rightarrow$  how much

a Real gas is compressible as compared to ideal gas under identical condition of  $(P, T, n)$

$$Z = \left( \frac{V_{\text{real}}}{V_{\text{ideal}}} \right)_{P, n, T}$$

# If  $Z > 1 \Rightarrow V_{\text{real}} > V_{\text{ideal}} \quad | \quad P, n, T$

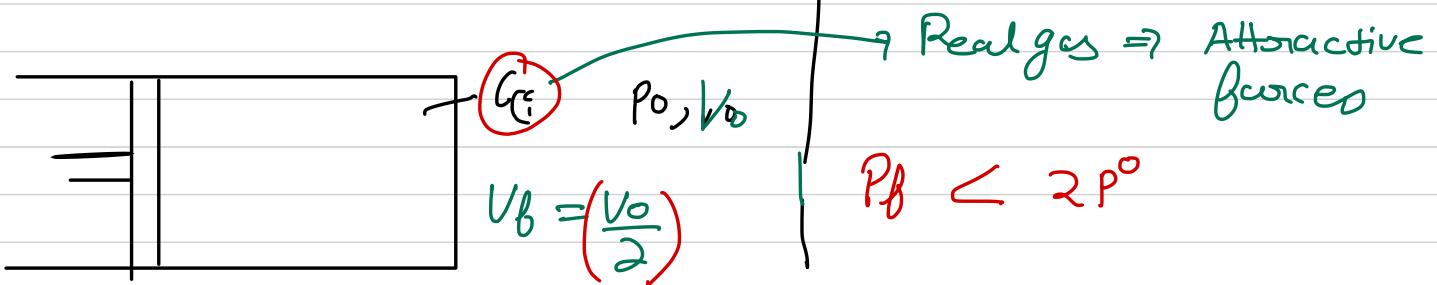
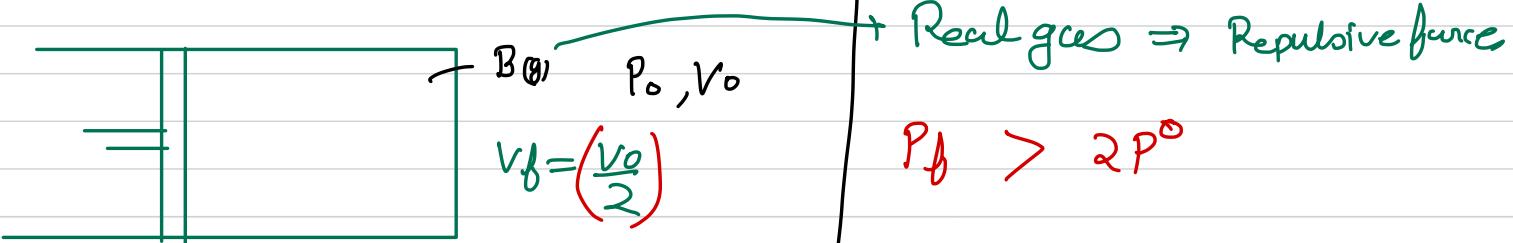
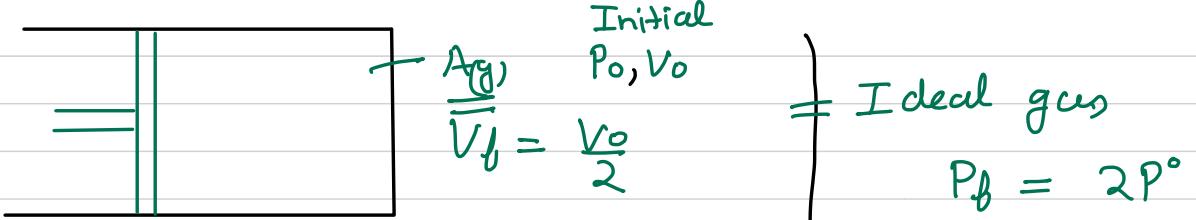
Repulsive forces  $\Leftarrow$  Less compressible

# If  $Z < 1 \Rightarrow V_{\text{real}} < V_{\text{ideal}} \quad | \quad P, n, T$

Attractive forces  $\Leftarrow$  More compressible

③  $Z = 1 \quad V_{\text{real}} = V_{\text{ideal gas}} \Rightarrow \begin{cases} \text{No attractive or} \\ \text{Repulsive forces} \end{cases}$   
(Ideal gas)

#



$$Z = \left( \frac{P_{\text{real}}}{P_{\text{ideal}}} \right) \underline{V, N, T}$$

$$Z = \left( \frac{V_{\text{real}}}{V_{\text{ideal}}} \right) \underline{P, N, T}$$

If we calculate the V<sub>molar</sub> of an Ideal gas

at 273K, 1 bar pressure, 1 mole

$$V_{\text{molar}} = \underline{\underline{22.7 \text{ Lits}}}$$

If Z > 1 for a gas, its Molar volume at 273, 1 bar will be V<sub>m</sub> > 22.7 Lits

If Z < 1 for a gas

$$V_m < 22.7 \text{ Lits}$$

# Equation of State for a Real gas.

Ideal gas Eq. of State

$$PV = nRT$$

Real gas  $\Rightarrow$  there are many equation of state

# We will only Derive and Study about a Real gas obeying an equation called vander-waal equation & such gases which obey Vander-waal eqn. are Vander-waal gases. #

$\Rightarrow$  Are all Real gases Vander-waal gases ?

NO

$\Rightarrow$  Are all Vander-waal gases, Real gases  
(Yes)

## Vander-waal gas

# Vander-waal during its derivation of eqn of state for a Real gas Assumed Only forces of Attraction to be present, he did not talk about the Repulsive forces & he took into account the molecular volume.

Vander-waal Real gas  $P_{\text{Real}} < P_{\text{ideal}}$

$P_{\text{Vander-waal gas}} > P_{\text{ideal}}$  X

Derivation of van der waal eqn.

$$\Rightarrow PV = nRT$$

$$P_{\text{ideal}} \times V_{\text{ideal}} = n \times R \times T$$

↓      ↓  
Ideal volume  
Ideal pressure

If we have a van der waal gas

$$P_{\text{r.w.g}} < P_{\text{ideal}}$$

\* For other Real gases  
'P' can be less or  
more than the  
ideal gas

so to get  $P_{\text{ideal}}$  of a V.W. gas we must add  
something to 'P' observed (Real) = V.W. gas , Let that be  
'b'

$$P_{\text{ideal ab V.W. gas}} = (P + b)$$

↓  
measured Pressure

$$\text{To get } V_{\text{ideal ab V.W. gas}} = (V_{\text{container}} - V_{\text{molecules}})$$

$$V_{\text{ideal}} = V - V_{\text{molecules}}$$

↓  
Container.

$$\overline{(P + b)(V - V_{\text{molecules}})} = nRT$$

$$b = \frac{an^2}{V^2}$$

= internal pressure  $\Rightarrow$

$n = \text{no. of molecules/gas}$

$a \Rightarrow$  van der waal  
const. for  
forces of  
attraction

$V = \text{measured volume}$   
 $= \text{Volume of container}$

(Ideal)



$P_{\text{ideal}}$

>

$\underline{P_{\text{real}}}$ )  $\Rightarrow$  measured pressure

$$* \quad \left( P + \frac{an^2}{V^2} \right) (V - V_{\text{molecules}}) = nRT$$

$$V_{\text{molecules}} \propto n_{\text{molecules}}$$

$$V_{\text{molecules}} = (n \times b) \quad b = \text{vander waal const.}$$

b = excluded volume per mole of gas

$$b = 4 \times \left( \frac{4}{3} \pi r^3 \right) \times N_A$$

↳ volume of 1 molecule assuming it to be a sphere

$$\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$$

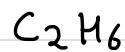
$$\left( P + \frac{a}{(V/n)^2} \right) \left( \frac{V}{n} - b \right) = RT$$

$$\boxed{\left( P + \frac{a}{V_m^2} \right) (V_m - b) = RT}$$

$\frac{V}{n} = V_m$   
↓  
molar volume

Vander-waal constant 'a'  $\Rightarrow$  For force of attraction

Gas	a (atm L <sup>2</sup> mol <sup>-2</sup> )	b (L mol <sup>-1</sup> )
Ammonia	4.17	0.0371
Argon	1.35	0.0322
Carbon dioxide	3.59	0.0427
Carbon monoxide	1.49	0.0399
Chlorine	6.49	0.0562
Ethane	5.49	0.0638
Ethanol	2.56	0.087
Ethylene	4.47	0.0571
Helium	0.034	0.0237
Hydrogen	0.024	0.0266
Hydrogen chloride	3.67	0.0408
Hydrogen bromide	4.45	0.0433
Methane	2.25	0.0428
Neon	0.21	0.0171
Nitric oxide	1.34	0.0279
Nitrogen	1.39	0.0319
Oxygen	1.36	0.0318
Sulphur dioxide	3.71	0.0564
Water	5.44	0.0305



Here 'a' only represents the weak Vander-waal

forces [i.e. induced dipole attractive force]

★ Size of molecules

more is the size of molecule more will

be contact surface area hence more will be forces (Vander-waal) of attraction

$$a(\text{O}_2) > a(\text{C}_2\text{H}_6) > a(\text{H}_2\text{O}) \\ > a(\text{NH}_3)$$

$$Z = \frac{V_{\text{real}}}{V_{\text{ideal}}} \Big| P, n, T$$

$$= \frac{P \times V_{\text{real}}}{(P \times V_{\text{ideal}})} \Big| n, T$$

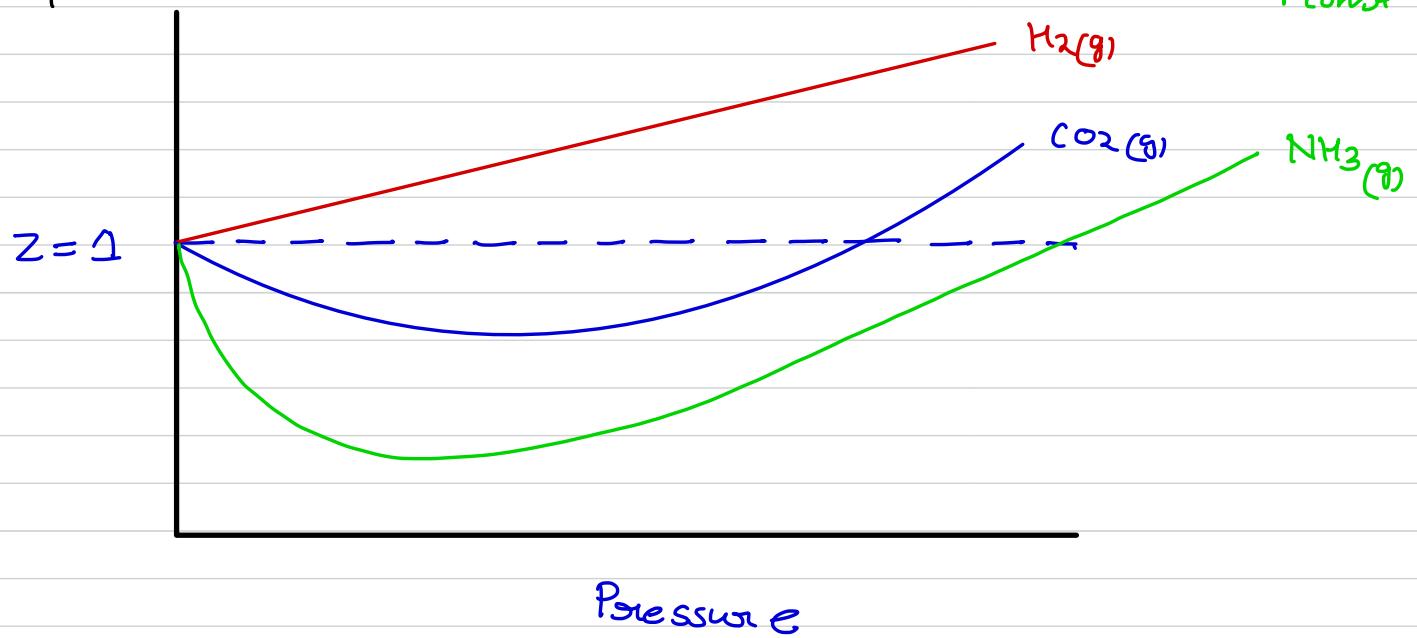
$$= \frac{P \times V_{\text{real}}}{n \times R \times T}$$

$$= \frac{P \times V_{\text{real}}}{R \times T} =$$

$$\boxed{\frac{P \cdot V_m}{R \cdot T} = Z}$$

volume of the real gas

Now we have an experimental graph plotted b/w ' $Z$ ' & 'P' for various gases at a given constant 'T'



Observation from the graph

- ① For Hydrogen gas  $Z > 1$  at all press.
- ② For all other gases at Low pressure  $Z < 1$
- ③ At high Pressure  $Z > 1$
- ④ At extremely low pressure  $Z \approx 1$  for all gases

$$\text{vanderwaal eq} \quad \left( P + \frac{a}{V_m^2} \right) (V_m - b) = RT \quad \approx$$

① **At Low Pressure**  $\rightarrow$  volume Large  $\rightarrow$  molecules are far apart from each other.  $\rightarrow$  attractive forces dominant.  $\&$  since volume is large we can NEGLECT the volume of molecules w.r.t. total volume of gas

$$V_m - b \approx V_m$$

i.e. 'b' can be neglected but NOT 'a'

$$\left( P + \frac{a}{V_m^2} \right) (V_m) = RT$$

$$PV_m + \frac{a}{V_m} = RT$$

$$PV_m = RT - \frac{a}{V_m}$$

$$\frac{PV_m}{RT} = 1 - \frac{a}{V_m RT}$$

$$Z = 1 - \frac{a}{V_m RT}$$

$$Z < 1$$

② **At high Pressure**  $\rightarrow$  volume small  
 $\rightarrow$  molecules are close to each other  $\rightarrow$   
 Repulsive forces will be dominant  $\rightarrow$   
~~'a'  $\rightarrow$  neglect~~, but 'b' can not be  
 neglected as the volume of gas is small.

$$(P)(V_m - b) = RT$$

$$PV_m - Pb = RT$$

$$PV_m = RT + Pb$$

$$\left(\frac{PV_m}{RT}\right) = 1 + \frac{Pb}{RT}$$

$$Z = 1 + \frac{Pb}{RT}$$

$$Z > 1$$

#3 At extremely Low Pressure

$$P \rightarrow 0 \Rightarrow V \rightarrow \infty \quad \frac{1}{V^2} \rightarrow 0$$

↓

$$\frac{V-nb}{V} \approx 1$$

$$\left(P + \alpha \times \frac{1}{V^2}\right) (V_m - b) = RT$$

$$P \times V_m = RT$$

$$Z = 1$$

\* For H<sub>2</sub> & He gas since molecular size is very-very small hence the force of attraction b/w these molecules are never dominant. i.e.

$\alpha \rightarrow 0$  neglected for H<sub>2</sub> & He

$$P(V_m - b) = RT$$

$$Z = 1 + \frac{Pb}{RT}$$

$$Z > 1$$

$$\left(P + \frac{a}{V_m^2}\right)(V_m - b) = RT$$

$$\Rightarrow P + \frac{a}{V_m^2} = \frac{RT}{V_m - b}$$

$$P = \frac{RT}{V_m - b} - \frac{a}{V_m^2}$$

$$PV_m = \frac{RT}{\left(1 - \frac{b}{V_m}\right)} - \frac{a}{V_m}$$

$$= RT \left(1 - \frac{b}{V_m}\right)^{-1} - \frac{a}{V_m}$$

$$= RT \left[1 + \frac{b}{V_m} + \frac{b^2}{V_m^2} + \frac{b^3}{V_m^3} + \dots\right] - \frac{a}{V_m}$$

$$= RT + \left(\frac{bRT}{V_m} - \frac{a}{V_m}\right) + \frac{b^2 RT}{V_m^2} + \frac{RTb^3}{V_m^3} + \dots$$

$$PV_m = RT + \frac{bRT - a}{V_m} + \frac{RTb^2}{V_m^2} + \frac{RTb^3}{V_m^3} + \dots$$

$$Z = 1 + \frac{b - a/RT}{V_m} + \frac{b^2}{V_m^2} + \frac{b^3}{V_m^3} + \dots$$

Virial equation of a Real Gas

If we take Low pressure condition.  $V_m$  is Large

we can neglect  $\frac{1}{V_m^2} \rightarrow 0$      $\frac{1}{V_m^3} \rightarrow 0$

$$Z = 1 + \frac{b - a/RT}{V_m}$$

If  $b - a/RT = 0$  i.e.  $(\text{Temp} = \frac{a}{Rb})$

$Z = 1$  at low pressure condition

↳ Ideal gas

↳ Boyle's Law obeyed

The Temperature at which a real gas behaves ideally over a low pressure zone is known as Boyle's Temp

at  $T = \frac{a}{Rb}$  gas is ideal  $Z = 1$

$$\left(b - \frac{a}{RT} > 0\right) \quad b > \frac{a}{RT} \Rightarrow \left(T > \frac{a}{Rb}\right)$$

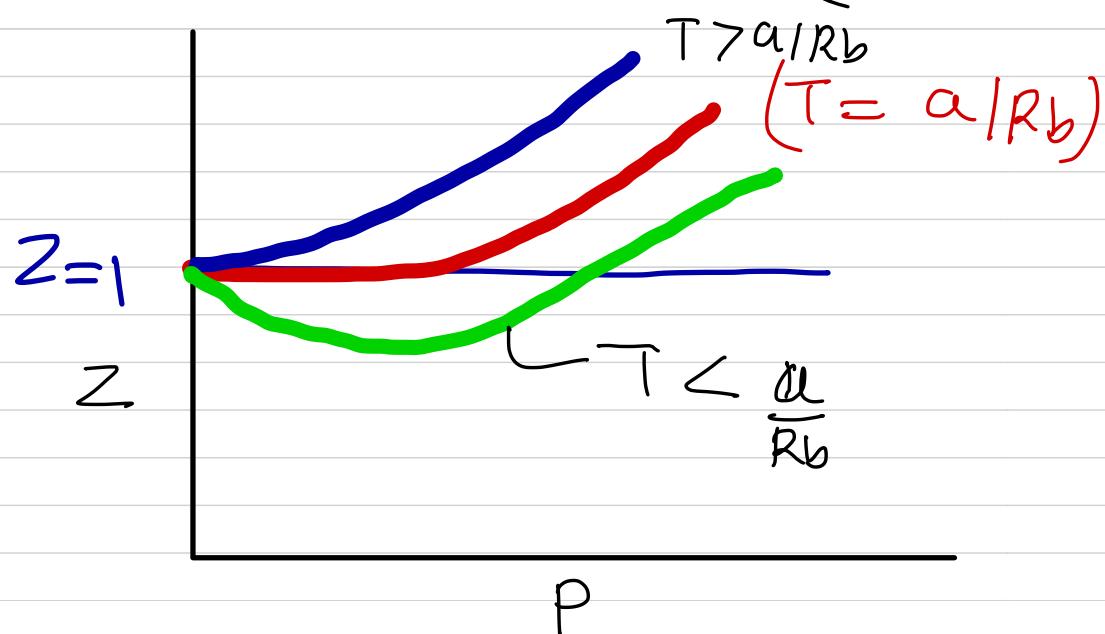
$$Z = 1 + \frac{\left(b - \frac{a}{RT}\right)}{V_m}$$

$Z > 1$  at Temp

$$\left(b - \frac{a}{RT} < 0\right) < 0$$

higher than Boyle's temp

$$Z < 1 \quad \left(T < \frac{a}{Rb}\right)$$

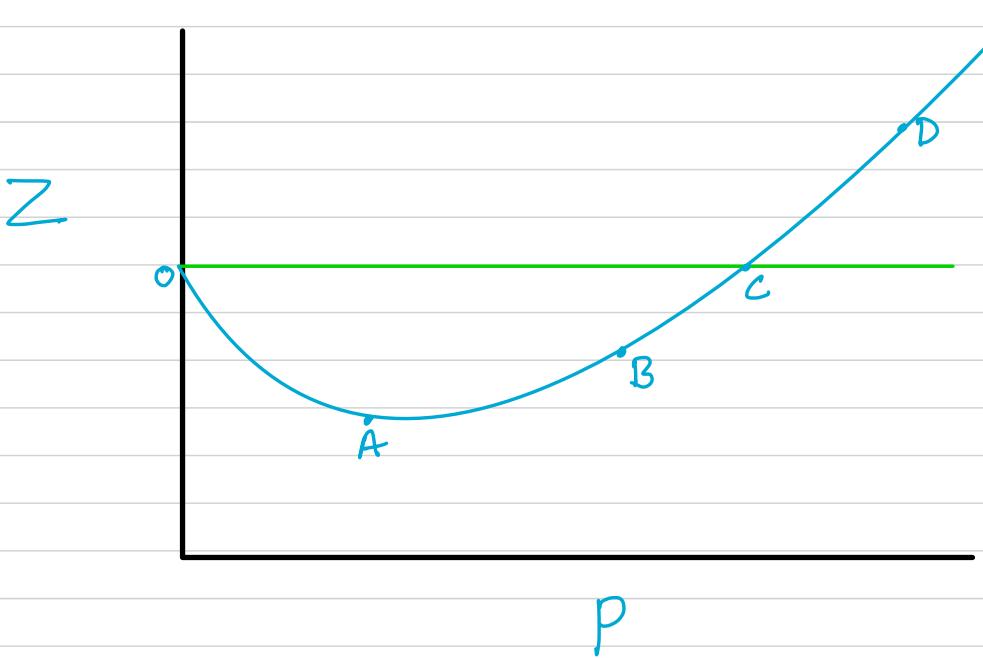


$$\textcircled{1} \quad \left(\frac{\delta Z}{\delta P}\right)_{P \rightarrow 0} = 0 \quad \text{at} \quad T = \frac{a}{Rb}$$

$$\textcircled{2} \quad \left(\frac{\delta Z}{\delta P}\right)_{P \rightarrow 0} > 0 \quad \text{at} \quad \left(T > \frac{a}{Rb}\right) \implies \text{Repulsive forces}$$

$$\textcircled{3} \quad \left(\frac{\delta Z}{\delta P}\right)_{P \rightarrow 0} < 0 \quad \text{at} \quad \left(T < \frac{a}{Rb}\right) \implies \text{Attractive force dominant}$$

$$\textcircled{4} \quad \left(\frac{\delta Z}{\delta P}\right)_{P \rightarrow \infty} > 0$$



$OA \rightarrow \left(\frac{\delta Z}{\delta P}\right) < 0$  attractive force dominant.

at 'A'  $\left(\frac{\delta Z}{\delta P}\right) = 0$  Maximum attractive forces

A  $\rightarrow$  B  $\rightarrow$  C  
 $\left(\frac{\delta Z}{\delta P}\right) > 0$  attractive forces dominant  
 $\cos Z < 1$

Magnitude of attractive force decreases and

Repulsive forces increases but dominance will be of Attractive

$C \rightarrow D$

$$\left(\frac{\delta z}{\delta P}\right) > 0$$

Repulsive forces dominant.

Calculate the volume occupied by 2 moles of a vanderwall gas at 5 atm(800 K)

Given :  $a = 4.0 \text{ atm } \ell^2 \text{ mol}^{-2}$  ,  $b = 0.0625 \ell \text{ mol}^{-1}$  ,  $R = 0.08 \ell \cdot \text{atm/K-mol}$

\*  $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$   $\times$  bcoz Temp is Boyle's Temp

\*  $T = 800K$   $a = 4.0$   $b = 0.0625$   $R = 0.08$

Find Boyle's Temp  $T_b = \frac{a}{R \times b} = \frac{4.0}{0.0625 \times 0.08} = 800K$

If gas is at Boyle's Temp then it behaves ideally.

$$PV = n \times R \times T \quad \checkmark$$

$$V = \frac{n \times R \times T}{P}$$

What is the compressibility factor ( $Z$ ) for 0.02 mole of a van der Waals' gas at pressure of 0.1 atm. Assume the size of gas molecules is negligible.

Given :  $RT = 20 \text{ L atm mol}^{-1}$  and  $a = 1000 \text{ atm L}^2 \text{ mol}^{-2}$

(A) 2

(B) 1

(C) 0.02

(D) 0.5

$$\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$$

$$\approx \left( P + \frac{an^2}{V^2} \right) V = nRT$$

$$PV + \frac{an^2}{V} = nRT$$

$$\frac{PV}{nRT} + \frac{an^2}{nRTV} = 1$$

$$\text{or } Z = 1 - \frac{an}{nRT}$$

same concept.

$$PV^2 + an^2 = nRTV$$

$$PV^2 - nRTV + an^2 = 0$$

$$\cancel{V^2} \quad \cancel{V} = \cancel{V}$$

If density of vapours of a substance of molar mass 18 g at 1 atm pressure and 500 K is 0.36 kg m<sup>-3</sup>, then calculate the value of Z for the vapours. (Take R = 0.082 L atm mole<sup>-1</sup> K<sup>-1</sup>)

\*

0.36 Kg gas has a volume of 1m<sup>3</sup>

18 gm gas has volume of  $\left( \frac{1m^3}{0.36 \times 1000} \times 18 \right)$

1 mole of gas has

Volume of  $\left( \frac{18}{0.36 \times 1000} m^3 \right)$

V<sub>m</sub>

$$Z = \left( \frac{PV_m}{RT} \right)$$

One litre gas at 400 K and 300 atm pressure is compressed to a pressure of 600 atm and 200 K. The compressibility factor is changed from 1.2 to 1.6 respectively. Calculate the final volume of the gas.

$$R_{\text{molar}} = \text{constant}$$

$$Z = \frac{V_{\text{real}}}{V_{\text{ideal}}} \mid P, n, T$$

$$= \frac{PV_{\text{real}}}{PV_{\text{ideal}}} = \frac{P \times V}{(n \times R) \times T} = Z$$

$$\frac{P_1 V_1}{Z_1 T_1} = \frac{P_2 V_2}{Z_2 T_2}$$

$$V = (1/3 \text{ Litre})$$

---

1 Derive the expression for compressibility factor of a vanderwaal gas at Boyle temperature.

v

-

The compression factor (compressibility factor) for one mole of a Van der Waals' gas at 0° C and 100 atmosphere pressure is found to be 0.5. Assuming that the volume of a gas molecule is negligible, calculate the Van der Waals' constant 'a'. **[JEE 2001]**

The density of the vapour of a substance at 1 atm pressure and 500 K is  $0.36 \text{ Kg m}^{-3}$ . The vapour effuses through a small hole at a rate of 1.33 times faster than oxygen under the same condition.

Determine

[JEE 2002]

- (i) mol. wt.;    (ii) molar volume;    (iii) compression factor  $z$  of the vapour and
- (iv) which forces among the gas molecules are dominating, the attractive or the repulsive

The density of water vapour at 328.4 atm and 800 K is 135.0 g/dm<sup>3</sup>. Determine the molar volume,  $V_m$  and the compression factor of water vapour .

If at 200 K & 500 atm, density of CH<sub>4</sub> is 0.246 gm/ml then its compressibility factor (Z) is approx  $2.0 \times 10^x$  . 'x' is:

The values of Van der Waals constant 'a' for the gases O<sub>2</sub>, N<sub>2</sub>, NH<sub>3</sub> and CH<sub>4</sub> are 1.360, 1.390, 4.170 and 2.253 L atm mol<sup>-2</sup> respectively. The gas which can most easily be liquefied is :

8. Consider the following table :

[JEE-MAINS-2019]

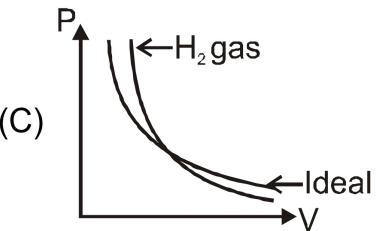
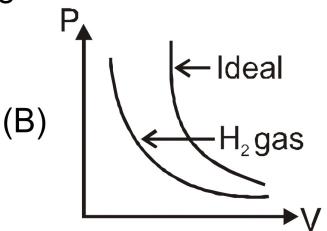
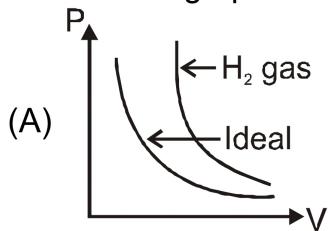
Gas	$a/(k \text{ Pa dm}^6 \text{ mol}^{-1})$	$b/(\text{dm}^3 \text{ mol}^{-1})$
A	642.32	0.05196
B	155.21	0.04136
C	431.91	0.05196
D	155.21	0.4382

a and b are vander waals constant. The correct statement about the gases is :

- (1) Gas C will occupy lesser volume than gas A; gas B will be lesser compressible than gas D
- (2) Gas C will occupy more volume than gas A; gas B will be lesser compressible than gas D
- (3) Gas C will occupy more volume than gas A; gas B will be more compressible than gas D
- (4) Gas C will occupy lesser volume than gas A; gas B will be more compressible than gas D

RC0046

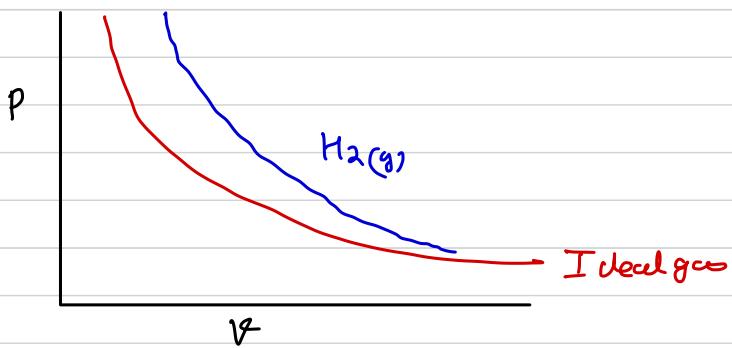
The correct graph to H<sub>2</sub> gas is:



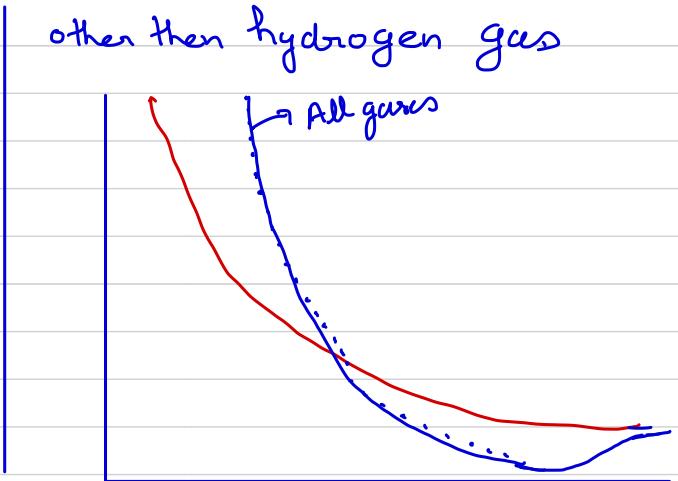
(D) All of these

concept for H<sub>2</sub> & He  $\gamma' > 1$

$$| V_{H_2 \text{ gas}} > V_{\text{Ideal}} \rangle_{P, N, T}$$



other than hydrogen gas



The critical density of the gas CO<sub>2</sub> is 0.44 g cm<sup>-3</sup> at a certain temperature. If r is the radius of the molecule, r<sup>3</sup> in cm<sup>3</sup> is: (N is Avogadro number)

(A)  $\frac{25}{\pi N}$

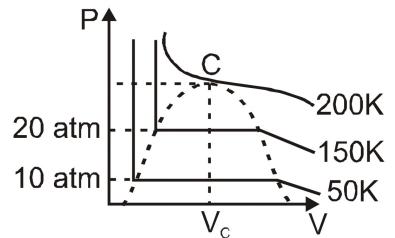
(B)  $\frac{100}{\pi N}$

(C)  $\frac{6}{\pi N}$

(D)  $\frac{25}{4N\pi}$

Identify the wrong statement related to the above graph :

- (A) between 50 K and 150 K temperature and pressure ranging from 10 atm to 20 atm matter may have liquid state.
- (B) zero is the maximum value of the slope of P-V Curve.
- (C) If vander waal equation of state is applicable above critical temperature then cubic equation of  $V_m$  will have one real and two imaginary roots.
- (D) At 100 K and pressure below 20 atm it has liquid state only.



If volume occupied by  $\text{CO}_2$  molecules is negligible, then calculate pressure  $\left(\frac{P}{5.277}\right)$  exerted by one mole of  $\text{CO}_2$  gas at 300 K. ( $a = 3.592 \text{ atm L}^2 \text{ mol}^{-2}$ ) assuming it follows

Vander waal equation.

$$\left(P + \frac{a}{V^2}\right) (V - b) = RT \quad "b" \text{ is negligible}"$$

$$\left(P + \frac{a}{V^2}\right) (V) = RT$$

$$PV + \frac{a}{V} = RT$$

$$T = 300\text{K}$$

$$a = 3.592$$

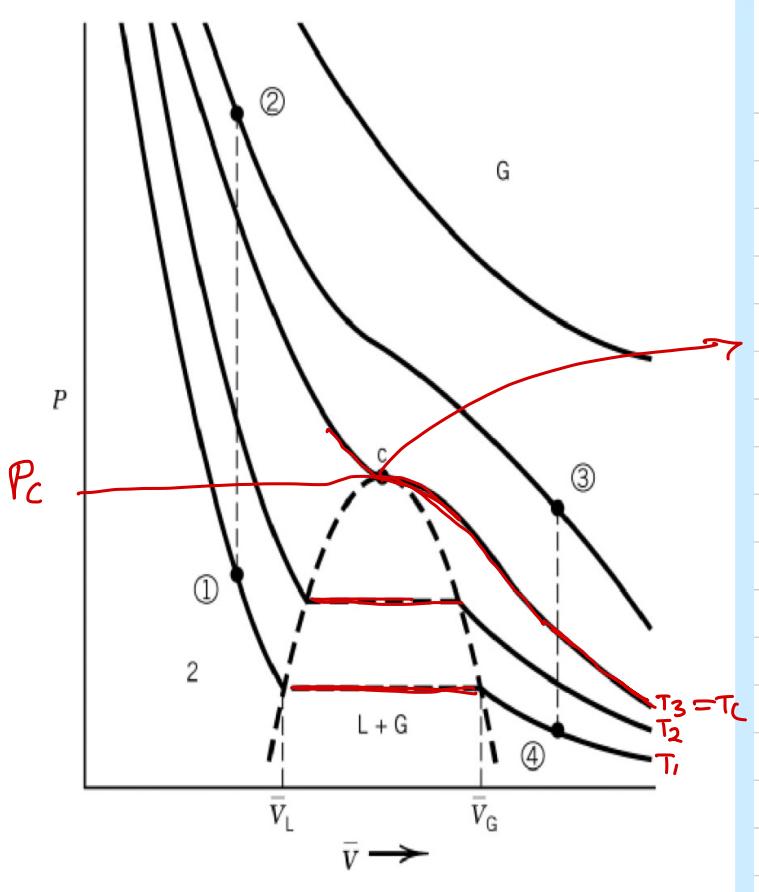
$$[PV^2 - RTV + a = 0]$$

\* At a given 'P' & T 'V' will be a single value.

$$D=0$$

$$R^2 T^2 - 4 \times Pa = 0$$

$$\boxed{\frac{R^2 T^2}{4a} = P}$$



curve inflection

$$\frac{\delta P}{\delta V} = 0 \quad \frac{\delta^2 P}{\delta V^2} = 0$$

Critical condition

## Concept      Liquification of Gases.

What can be done to Liquify the gas

# Molecules must be brought close to each other.

1) Increase the Pressure      ② Temperature decrease

# But if there are NO ATTRACTIVE FORCES b/w gas molecules, then they will again move apart that is it won't be liquified  $\Rightarrow$

Ideal gases cannot be Liquified

"Real gases can be liquified"

# If Temperature of Gas is higher than a certain value then the Kinetic energy of gas molecules will be high, so even on applying pressure (Howsoever large) the gas will no liquify.

It implies that to Liquify a gas it must be first cooled below a certain Temperature.,

The Temperature above which a gas can not be Liquified by only applying pressure is Known as Critical Temperature.

The pressure needed to Liquify the gas at Critical Temperature is known as Critical pressure.

volume occupied by 1 mole of gas at critical pressure & Temp is known as critical volume.

$$P_c = \frac{a}{27b^2}$$

Critical Press

$$T_c = \frac{8a}{27Rb}$$

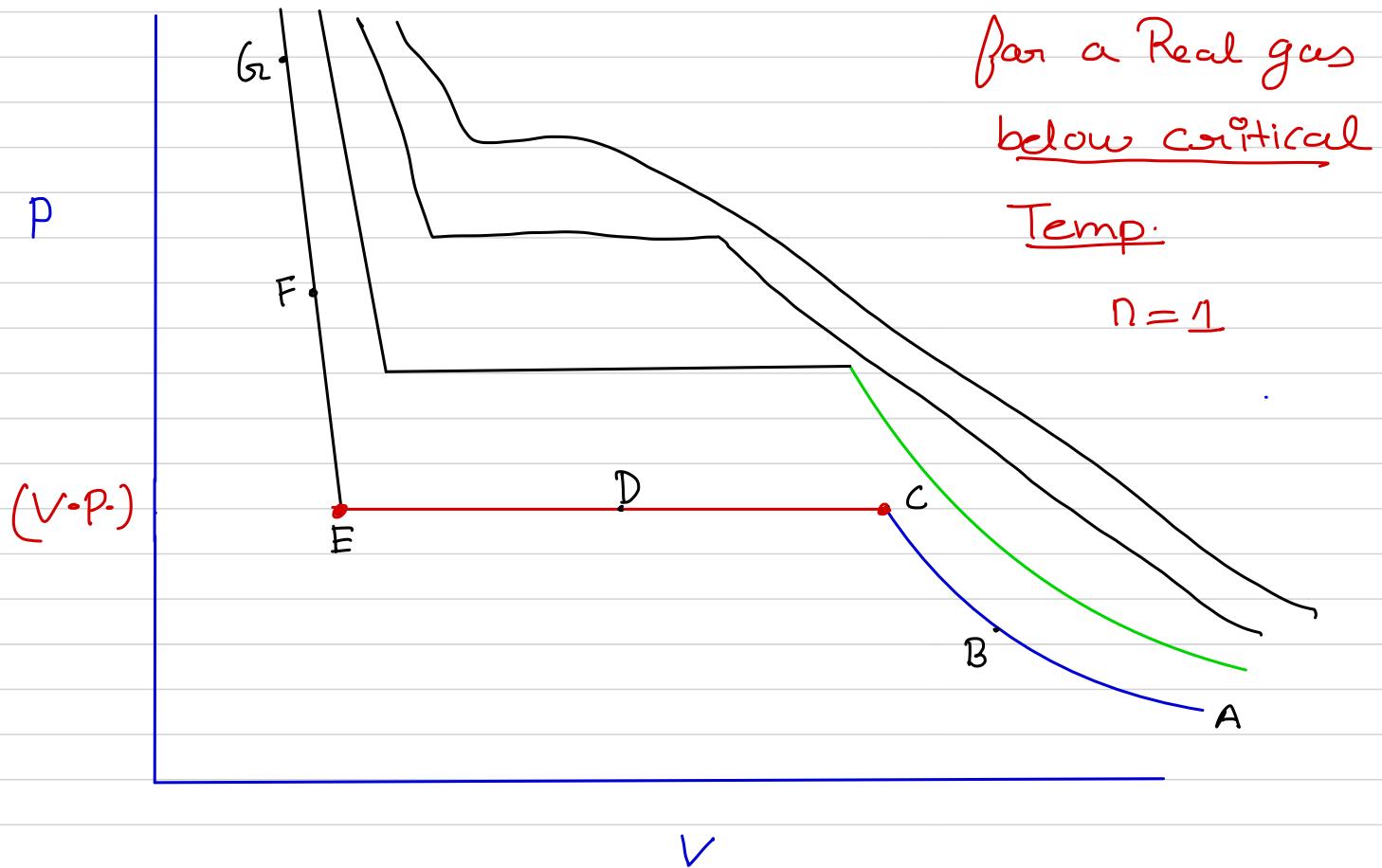
Critical Temp

$$V_c = 3 \times b$$

Critical volume.

where 'a' & 'b' are van der waal constant as in vander waal egn.

### Andrew's Isotherms for Real gas



From A  $\rightarrow$  B  $\rightarrow$  C Gas is getting compressed at constant Temp so its volume is decreasing then its Density is increasing.

$C \rightarrow$  Liquefaction starts

$\& C \rightarrow D \rightarrow E \rightarrow$  Gas is getting liquefied at constant 'P'

  
Gas & Liquid are in equilibrium  
→ Vapor pressure.

(E) Complete Liquid

$E \rightarrow F \rightarrow G \rightarrow$  compressing a Liquid.

At critical point there is no distinction b/w  
(Liq & Gas) it is known as "Fluid"

$$T_b = \frac{a}{Rb} \quad T_c = \frac{8a}{27Rb}$$

$$T_c = \frac{8}{27} \times T_b$$

$$Z_c = \frac{P_c \times V_c}{T_c \times R} = \frac{3}{8}$$