GASEOUS STATE

Matter → Solid / liquid / Gas

	Solid	liquid	Gas
Intermolecular force	Large	small	almost zero
Density	High	Intermediate	Low
Volume	Fixed	Fixed	Varies
Geometry	Definite	not-definite	not-definite
Thermal energy	almost zero	greater	High
Compressibility	not compressible	not easily compressible	Compressible

→ Parameters to define properties of Gas →

① Pressure →

$P = \frac{Force}{Area}$ N/m^2 or Pa - SI, Dyne - CGS

other Unit → atm

1 atm → 1.013×10^5 Pa
→ 1.013 bar
→ 760 torr
→ 760 mm of Hg
→ 76 cm of Hg

(Vimp)

② Volume →

SI Unit = m^3 ✓

$1\ m^3 = 1000\ dm^3 = 1000\ ml$ ✓

$1\ dm^3 = 1\ litre = 1000\ ml = 1000\ cm^3 = 1000\ cc$ ✓

③ Temperature →

SI Unit = Kelvin ✓

$T_K = °C + 273$

$\frac{C}{5} = \frac{F-32}{9}$

④ Amount of Substance →

(→ mole (unit))

→ Thermal Energy

→ Directly proportional to Heat.

→ Temp ↑, random motion ↑, K.E. of particle ↑, Thermal energy ↑

→ Intermolecular force →
Keeps molecules together.

→ Thermal Energy →
Keeps molecules apart.

Solid liq. Gas → Intermolecular force ↓
→ Thermal energy ↑.

← Gaseous State →

→ 11 elements are gases. (*)

H, He, N, O, F, Ne, Cl, Ar, Kr, Xe, Rn

→ Physical properties of Gas →

→ Highly Compressible
→ Exert pressure equally in all direction
→ Low density than solid/liquid.
→ Vol./Shape not fixed.
→ Intermix evenly and completely (Homogenous mixture)

Gas-laws →

① Boyle's Laws →

T → Constant, n → constant

$PV = nRT$
(R → Constant)

$P \propto \frac{1}{V}$

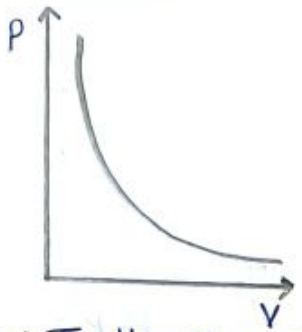
$P = K\frac{1}{V}$ → Depends on temperature and moles of gas.

$PV = K$

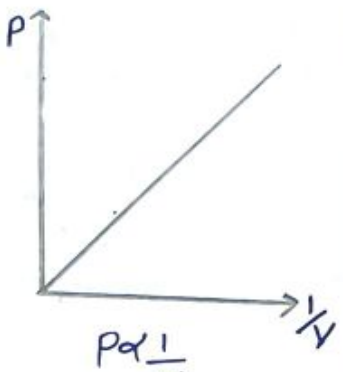
$P_1V_1 = P_2V_2$ ✓ ✓

$V \propto T$ (*)

NEET SLAYER

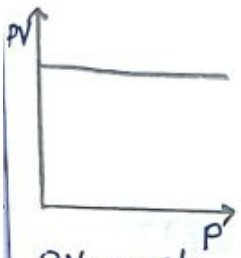


→ Isotherm graph
→ PV = Constant
→ xy - Constant

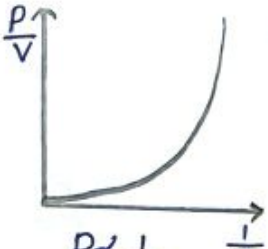


$P \propto \frac{1}{V}$

$y \propto x$



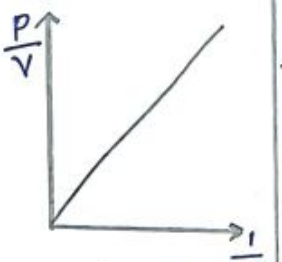
PV = const.
y = const.



$P \propto \frac{1}{V}$

$\frac{P}{V} \propto \frac{1}{V^2}$

$y \propto x^2$



$\frac{P}{V} \propto \frac{1}{V^2}$

$y \propto x$

(Que) Balloon is filled with Hydrogen at room temp. It will burn if pressure exceeds 0.2 bar. If at 1 bar pressure, gas occupies 2.27 litres Volume. Max Vol. of balloon?

→ Temp → Constant.

$P_1V_1 = P_2V_2$

$V_2 = \frac{P_1V_1}{P_2}$

$= \frac{1 \text{ bar} \times 2.27 \text{ l}}{0.2 \text{ bar}} = 11.35 \text{ litre}$

Charle's Law

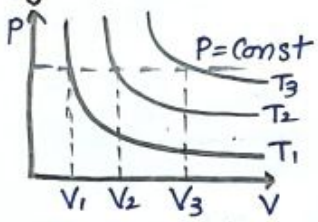
pressure → Const, Moles → Const

$V \propto T_{(K)}$

$V = KT_{(K)}$

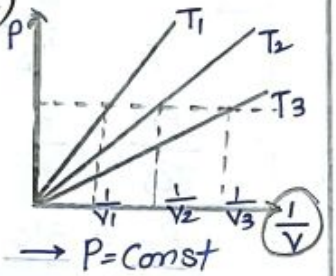
$\frac{V_1}{V_2} = \frac{T_1}{T_2}$ $\quad$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

(Que)



→ $V_3 > V_2 > V_1$
→ P = Constant
→ $V \propto T$
→ $T_3 > T_2 > T_1$

(Que)

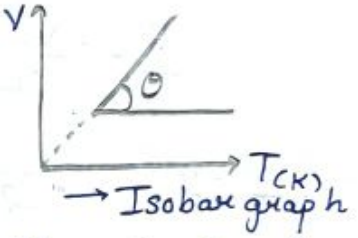


→ P = Const
→ $V \propto T$
→ $\frac{1}{V_3} > \frac{1}{V_2} > \frac{1}{V_1}$
→ $V_3 < V_2 < V_1$
→ $T_3 < T_2 < T_1$

$PV = nRT$

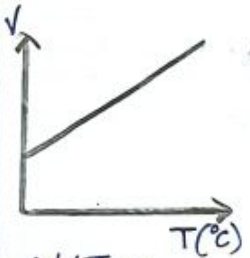
$V = \frac{nR}{P}T$

$V = KT_{(K)}$



→ Isobar graph

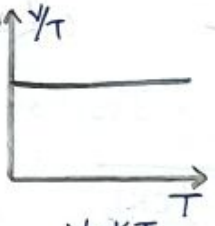
Slope = $\tan\theta = \frac{nR}{P}$



$V \propto T_{(K)}$

$V \propto (273 + t^\circ C)$

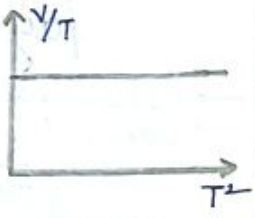
$y \propto c + x$



$V = KT$

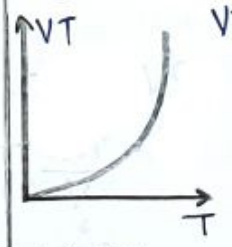
$\frac{V}{T} = K$

$y = K$



$V = KT$

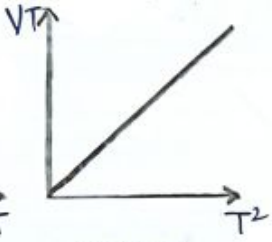
$\frac{V}{T} = K$



$V = KT$

$VT = KT^2$

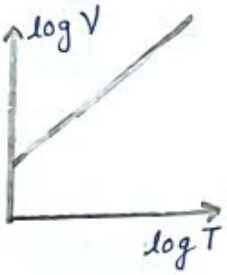
$y = Kx^2$



$V = KT$

$VT = KT^2$

$y = Kx$



$V = KT$

$\log V = \log K + \log T$

$y = c + x$

Relation b/w density and pressure using Boyle's Law →

$PV = K$; $d = \frac{m}{V}$ → $V = \frac{m}{d}$

$P \times \frac{m}{d} = K$

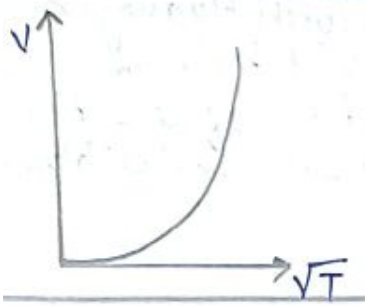
$P = \frac{K}{m} \cdot d$

$P = K'd$

$P \propto d$ → at high pressure, gases are more denser

(*)

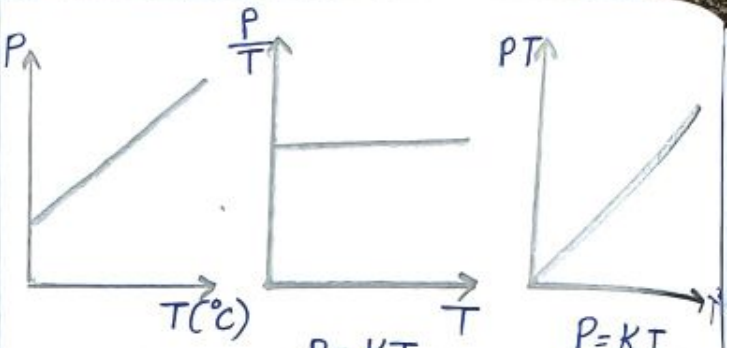
NEET SLAYER



$V = KT$

$y = K(\sqrt{T})^2$

$y = Kx^2$



$P = KT$

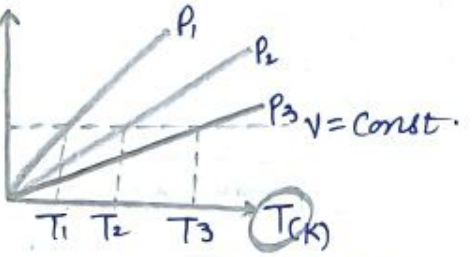
$\frac{P}{T} = K$

$P = KT$

$PT = KT^2$

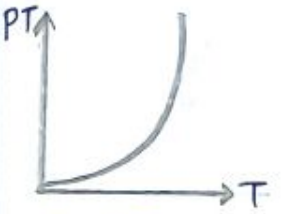
$y = x$

(Que)



$V = Const.$ $\quad T_3 > T_2 > T_1$

$P \propto T$ $\quad P_3 > P_2 > P_1$



$P = KT$

$PT = KT^2$

$y = x^2$

$P = KT$

$P = K(\sqrt{T})^2$

$y = Kx^2$

Absolute Zero ⟶ Lowest hypothetical/imaginary temp. at which gases suppose to occupy zero Vol. (−273°C / 0K)

(Que) At temp. 23.4°C, Vol. of balloon is 2 litre. What will be final vol. of balloon, when temp. rises to 26.1°C.

⟶ $P = Const.$ $\quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$V_2 = \frac{V_1}{T_1} \times T_2 = \left(\frac{2}{296.4} \times 299.1\right)$$

$$= 2.018 \text{ litre}$$

(Que) Temp. of gas was increased from 27°C ⟶ 37°C at const. Vol. What will be pressure of Gas?

⟶ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$.

$\frac{P_1}{300} = \frac{P_2}{310}$

$\frac{P_2}{P_1} = \frac{310}{300}$ ⟶ $P_2 = \frac{31}{30} P_1$

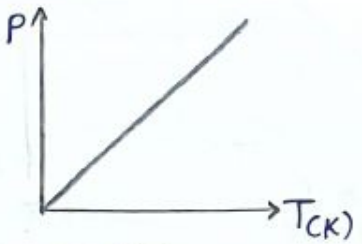
Gay-lussac Law ⟶

⟶ Const Vol
⟶ Const. Mol } Isochor Graph

$P \propto T_{(K)}$

$P = KT$

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $\quad \frac{P_1}{P_2} = \frac{T_1}{T_2}$



$P = KT$

$y = Kx$

$PV = nRT$

$P = \frac{nR}{V} T$

slope

Avogadro Law

Pressure ⟶ Const., Temp ⟶ Const.

$V \propto n$

⟶ equal Vol. of all gases, contain equal no. of molecule and not atom

$V = Kn$

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$ $\quad \frac{V_1}{V_2} = \frac{n_1}{n_2}$.

$V \propto n$

$V = Kn$

$V = K\frac{m}{M}$

$M = K\frac{m}{V}$

$M = Kd$

Ideal Gas Equation

(1) Boyle's Law	Temp. Const.	$V \propto \frac{1}{P}$	$PV = \text{Const.}$	$P_1V_1 = P_2V_2$
(2) Charle's Law	Pressure Const.	$V \propto T$	$V = KT$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
(3) Gay Lussac Law	Vol. Const.	$P \propto T$	$P = KT$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
(4) Avogadro Law	Pressure & Temp. Const.	$V \propto n$	$V = Kn$	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Ideal Gas → Gas that follows all these laws (Hypothetical) ✓

Combining all Gas Laws.

$V \propto n\frac{T}{P}$

$V = R\frac{nT}{P}$

$PV = nRT$ → Universal Constant

$R = 8.314\ JK^{-1}mol^{-1}$
$R = 0.0821\ \text{atm-litre}\ K^{-1}mol^{-1}$
$R = 1.98\ \text{cal}\ K^{-1}mol^{-1}$

imp

Combined Gas Law

$PV = nRT$

$\frac{PV}{T} = nR \rightarrow \text{constant}$

$\frac{PV}{T} = \text{Constant}$

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \text{constant}$.

Que) $P_1 = 760\ mmHg$, $P_2 = ?$
$V_1 = 600\ ml$, $V_2 = 640\ ml$
$T_1 = 298K$, $T_2 = 283\ K$

Sol^n:→ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$P_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{V_2}$

$= 676.6\ mmHg$

Relationship b/w molar mass and density →

$PV = nRT$

$PV = \frac{m}{M}RT$

$PM = \frac{m}{V}RT$

$PM = dRT$

$M = \frac{dRT}{P}$ ✓✓

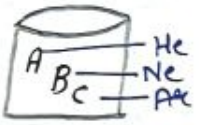
NEET SLAYER

Dalton's Law of Partial Pressure

→ Valid only for non-reactive gases

→ Total pressure of mixture is sum of individual (partial) pressure.

$P_T = P_A + P_B + P_C$ ✓



$P_A = \frac{n_ART}{V}$ $P_B = \frac{n_BRT}{V}$ $P_C = \frac{n_CRT}{V}$

$P_T = (n_A + n_B + n_C)\frac{RT}{V}$

$\frac{P_A}{P_T} = \frac{n_A}{n_A+n_B+n_C}$, $\frac{P_B}{P_T} = \frac{n_B}{n_A+n_B+n_C}$, ✓

$\frac{P_C}{P_T} = \frac{n_C}{n_A+n_B+n_C}$. ✓

$\frac{P_A}{P_T} = \chi_A = P_A = \chi_A \cdot P_T$ ✓

$\frac{P_B}{P_T} = \chi_B = P_B = \chi_B \cdot P_T$ ✓

$\frac{P_C}{P_T} = \chi_C = P_C = \chi_C \cdot P_T$ ✓

Que) A ($Ne + O_2$) mixture contain 70.6 gm O_2 and 167.5 gm Ne. If total pressure is 25 bar. Find partial pressure of Ne and O_2.

Sol^n → $Ne + O_2$

$n = \frac{167.5}{20}$ $\quad$ $n = \frac{70.6}{32}$

$n = 8.37$ $\quad$ $n = 2.21$

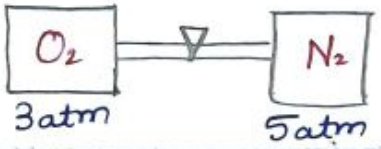
$P_{O_2} = \chi_{O_2} \cdot P_T = \frac{2.21}{2.21+8.37} \times 25\,bar$

$P_{Ne} = \chi_{Ne} \cdot P_T = \frac{8.37}{2.21+8.37} \times 25\,bar$

Gas Collected Over Water

$P_{dry\ gas} = P_{total} - aq.\ tension$ → Pressure of water vapour

Graham's Law Of Diffusion



→ Stopper removed

↓

Diffusion takes place

Graham's Law of Diffusion →

Constant Temp. and pressure

→ Rate of diffusion is inversely proportional to sq. root of density

$r \propto \frac{1}{\sqrt{d}} \rightarrow \frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}}$

$M = \frac{dRT}{P} \rightarrow \frac{\sqrt{M_2}}{\sqrt{M_1}} \rightarrow VD = \frac{M}{2} \rightarrow \frac{\sqrt{VD_2}}{\sqrt{VD_1}}$

Kinetic Theory Of Gas

→ Postulates →

1. ~~(1)~~ Gases consist of very small atoms/molecules whose volume is negligible compared to container's Volume.
2. ~~(2)~~ There is no interaction b/w gas particles.
3. Gaseous molecules are under random motion.
4. Due to random motion, gas particles collide with wall of container and hence exert pressure
5. Collision of gas molecule is perfectly elastic. → Total energy of molecule, before and after collision remains same
6. Molecules move with different speed, however speed of each molecule keeps on changing as collission occurs.
7. Avg. Kinetic energy of system depends on temp.

Kinetic Gas Equation

$PV = \frac{1}{3} m N V^2_{rms}$ ★

Que) Container of Capacity 1 litre have 10^{23} molecule each having mass 10^{-22} gram. If rms speed is 10^3 m/s. Calculate pressure of Gas in Pascal

Sol^n → $PV = \frac{1}{3} m N V^2_{rms}$

1 litre $= 10^{-3}\ m^3$

$P = \frac{\frac{1}{3} \times 10^{-25} \times 10^{23} \times 10^6}{10^{-3}}$

$= \frac{1}{3} \times 10^7$ pascal

Average Kinetic Energy

$PV = \frac{1}{3} m N V^2_{rms}$

$PV = \frac{2}{3}\left(\frac{1}{2} m N V^2_{rms}\right)$

$PV = \frac{2}{3} KE.$

$KE = \frac{3}{2} PV = \frac{3}{2} nRT$ ★ → avg. Kinetic energy

→ avg. Kinetic energy of 1 mole

$KE = \frac{3}{2} RT$ ★

→ avg. Kinetic energy of 1 molecule

$KE = \frac{3}{2} \frac{RT}{N}$ ★

$KE = \frac{3}{2}KT$ → $K = \frac{R}{N_A}$ = Boltzman Constant

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

Molecular Speed

① RMS Speed →

$V_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3KT}{m}}$

② Avg. Speed →

$V_{avg} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$

④ Most Probable Speed →

$V_{mp} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2KT}{m}}$

$R > A > M$

$V_{rms} > V_{avg} > V_{mp}$

Ratio →

$V_{rms} : V_{avg} : V_{mp} = \sqrt{3} : \sqrt{\frac{8}{\pi}} : \sqrt{2}$

Real Gas

↳ Gases which obey gas laws only at low pressure and high temperature.

Real Gas deviates from ideal behaviour, because →

① Real gas have finite volume.

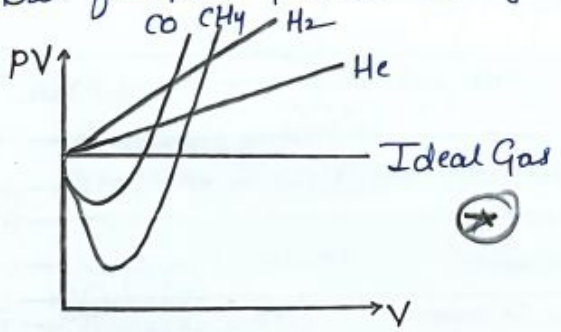
② Intermolecular interaction is not zero.

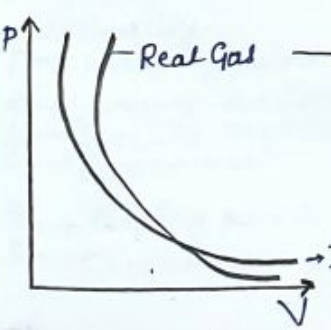
Deviation of Real Gas

Boyle's Law → PV Const.

PV V/s P Graph → ||straight.

But for Real Gas, not straight.





→ At high pressure, real gases have more vol. than ideal gases.

→ At low pressure, both Vol are same

At High Pressure →

→ Molecules of gases are close to each other.

→ Molecular attractive forces, starts operating

→ Molecules do not strike the walls of container, with full impact, because of attractive force by other molecule

→ pressure exerted by real gas, is less than ideal gas.

→ $P_{ideal} = P_{real} + \frac{an^2}{V^2}$ } correction term

At high Pressure →

Vol of gas molecule also become significant

→ Vol of gas molecule = nb

Vanderwall Equation →

$PV = nRT$ → Ideal Gas eq.n

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

↳ a, b → Vanderwall Constant.

a

mag. of intermolecular attractive forces.

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

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

$a = \frac{PXV^2}{n^2} = \frac{(N/m^2)l^2}{(mol)^2}$

$\frac{atm\ l^2}{(mol)^2}$

b

Size of gas molecule.

$(V - nb)$

nb = Volume

$b = \frac{Vol.}{n}$

= litre/mol

Compressibility Factor

Deviation from ideal gas is explained by compressibility factor.

$Z = \frac{(PV)_{real}}{(PV)_{ideal}} = \frac{(PV)_{real}}{nRT}$

$Z = 1$ → Ideal Gas

$Z \neq 1$ → Real Gas

$Z > 1$ → +ve deviation

$Z < 1$ → -ve deviation

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

$$P_cV_c=\frac{3}{8}RT_c$$

→ One Mole:

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

$$PV-Pb+\frac{a}{V}-\frac{ab}{V^2}\to 0=RT$$

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

$$\frac{PV}{RT}=1+\frac{Pb}{RT}\to 0-\frac{a}{VRT}$$

at low pressure →

$$\frac{PV}{RT}=1-\frac{a}{VRT}$$

$$Z=1-\frac{a}{VRT}$$

$Z<1$

-ve deviation

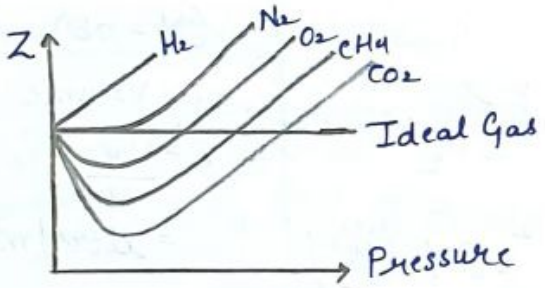
at High pressure →

$$\frac{PV}{RT}=1+\frac{Pb}{RT}-\frac{a}{VRT}\to 0$$

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

$Z>1$

+ve deviation



Boyle Temp. → Temp. at which real gas obeys ideal gas laws, over a range of temp.

Liquification Of Gas

Gas → Liquid → Pressure ↑↑, → Temp ↓↓ (Cooling } Both

Critical Temperature →

Temp. above which, gas can't be liquified.
Corresponding pressure → Critical pressure
Corresponding Value → Critical Volume.

$$T_c=\frac{8a}{27Rb},\quad P_c=\frac{a}{27b^2},\quad V_c=3b$$