Saturday

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states of matter

GASEOUS STATE.

Intermolecular force (Vanderwaal's):

- the intermolecular force of attraction between two identical molecules or two different molecules known as VANDERWAAL'S force.

- Vanderwaal forces are responsible for physical properties (such as Boiling point (BP), Melting point (MP) solubility, physical state, etc.) of chemical substance

- following are the different types of intermolecular forces (IMF):

a) Polar - Polar IMF or Dipole - Dipole IMF.
b) Polar - Non polar IMF or Dipole - Induced dipole IMF.
c) Non polar - Non polar IMF or Induced dipole - Induce dipole IMF or london force or Dispersion force.
d) Ion - Polar IMF or Ion - Dipole IMF.
e) Ion - Non Polar IMF or Ion - Induced Dipole IMF.

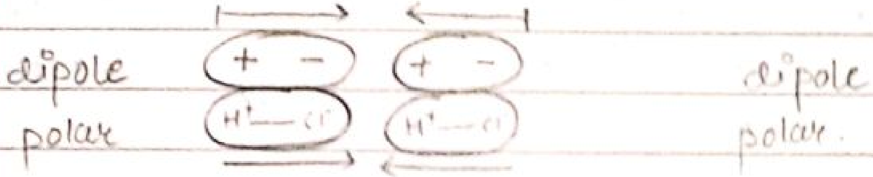
- IMF is not a bond it is just a force of attraction.

Polarity ∝ IMF

(A.) # Polar-Polar IMF:

- force of attraction carried out between two polar molecules.

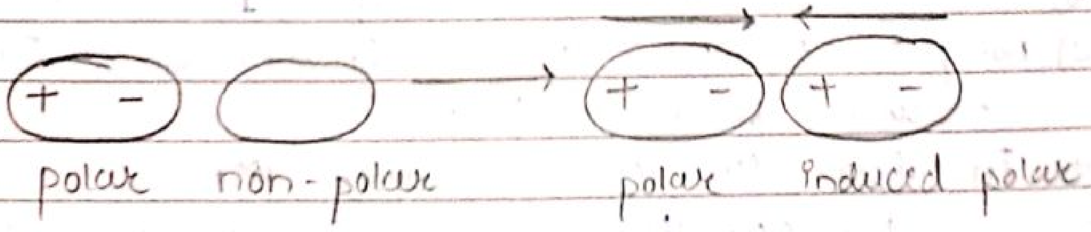
eg: HCl, HBr, H_2O (molecules). (force of attraction) etc.



(B.) # Polar- Non polar IMF:

- intermolecular force of attraction carried out between polar and non-polar molecules.

eg: Polar molecules (HCl, HBr, etc.) induce polarity in Non-polar (O_2, F_2, Ar, Kr, etc.)



(C.) # Non-polar - Non-polar IMF:

- weakest I.M.F, carried out between two non-plolar molecules.

eg: IMF between inert gases O_2, N_2, CO_2, etc.

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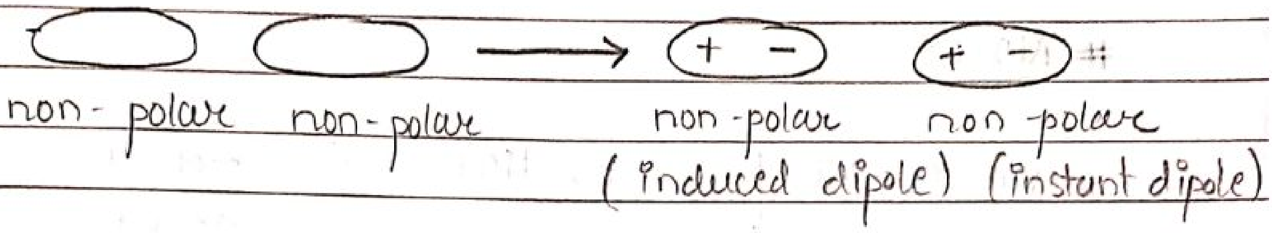
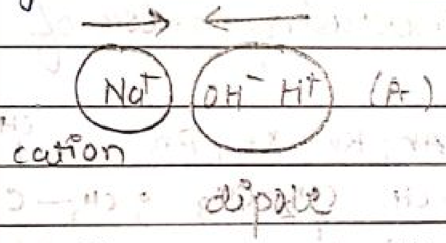


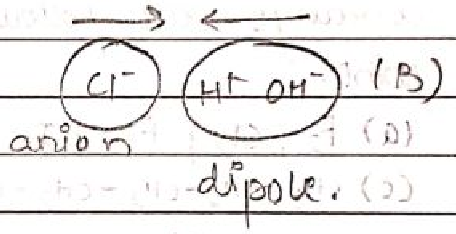
$\boxed{IMF \propto \begin{array}{l} at.wt. \\ mol.wt. \end{array}}$

(D) # Ion-Dipole IMF:

- IMF carried out between ion (cation/anion) and polar molecules.

eg: interaction between (a) Na^+ H_2O (b) Cl^- H_2O.

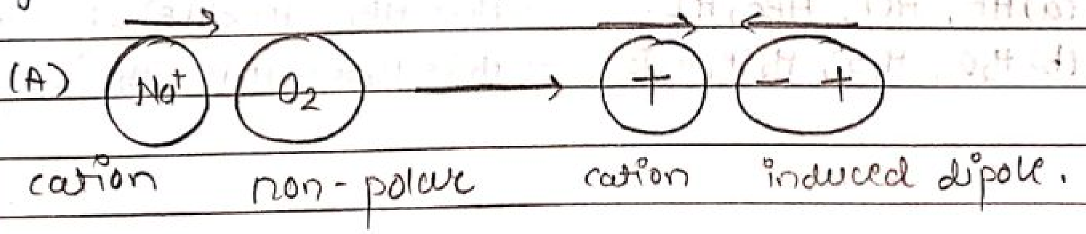


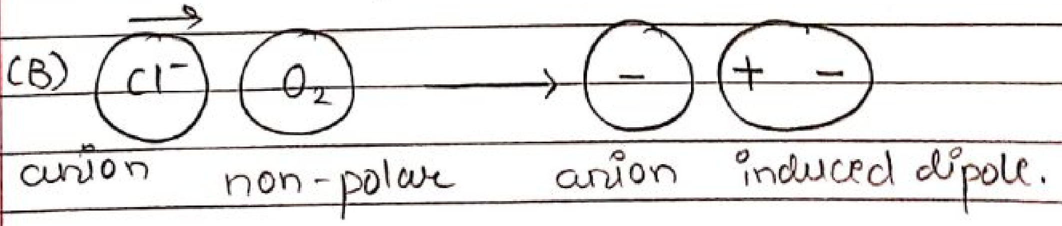


(E.) # Ion-Induced dipole IMF:

- IMF carried out between ion (cation/anion) and non-polar molecules.

eg: (A) $Na^+ + O_2$ (B) $Cl^- + O_2$





(III) 2-2 dimethyl propane ; neopentane

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NOTE:

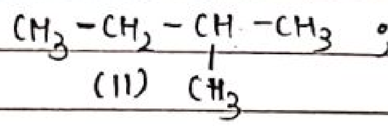
- Vanderwall force or IMF $\propto \frac{\text{mol. wt}}{\text{ato. wt.}}$
-
- Vanderwall force or IMF $\propto$ surface area
- Vanderwall force or IMF $\propto$: Boiling point

(up to down BP↑)

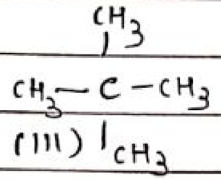
Q:1 arrange the following, in increasing order of boiling point.

(a) F_2 ; Cl_2 ; Br_2 ; I_2 (b) Ar ; Kr ; Xe ; Rn

(c) $CH_3-CH_2-CH_2-CH_2-CH_3$ (I) ;



$CH_3-CH_2-CH(CH_3)-CH_3$ (II) ;



$CH_3-C(CH_3)_2-CH_3$ (III)

→ (a) $F_2 < Cl_2 < Br_2 < I_2$ (VWF↑)

(b) Ar < Kr < Xe < Rn (VWF↑)

(c) (III) < (II) < (I) (SFA↑)

Q:2 arrange the following, in increasing order of boiling point.

(a) HF, HCl, HBr, HI. → HCl < HBr < HI < HF ‿ hydrogen bond.

(b) H_2O, H_2S, H_2Se, H_2Te → $H_2S < H_2Se < H_2Te < H_2O$ ‿ hydrogen bond.

Measurable properties of gases:

ideal gas situ. = 0 f. of attra.

a) Mass: gm, mole.

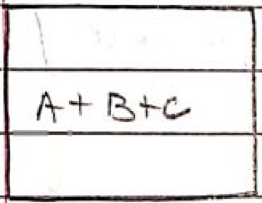
b) Temperature: always Kelvin ; Temp↑ K.E↑

c) Volume: $1\,ltr = 1000\,ml = 1\,dm^3 = 1000\,cc\,(cm^3) = 10^{-3}\,m^3$

→ volume of any gas = volume of container.

→ if 3 numbers of non-reactive gases available in a container

→ volume of each gas = volume of container.



Volume of Gas A $= 100\,cm^3$

volume of Gas B $= 100\,cm^3$

Volume of Gas C $= 100\,cm^3$

⟹ vol. of container $= 100\,cm^3$

d) Pressure: pressure of gas is defined as the force exerted by the gas on the wall of container.

→ it is isotropic in nature, due to its same presence in all three dimensions

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}} \longrightarrow \frac{m \cdot g}{A} \longrightarrow \frac{\overset{\text{volume}}{v} \times \overset{\text{density}}{d} \times g}{A}$$

$\rho = \frac{m}{V}$

∴, $m = d \cdot v$

∵ vol [illegible] h × area

$$\Longrightarrow \frac{\overset{\text{area}}{\cancel{A}} \times \overset{\text{height}}{h} \times d \times g}{\cancel{A}}$$

$$\boxed{P = hdg} \quad \text{or} \quad \boxed{P_{gm} = \rho hg}$$

v = volume

d or ρ = density

h = height

A = area

g = gravitational acceleration $9.8\,m/s^2$ or $10\,m/s^2$.

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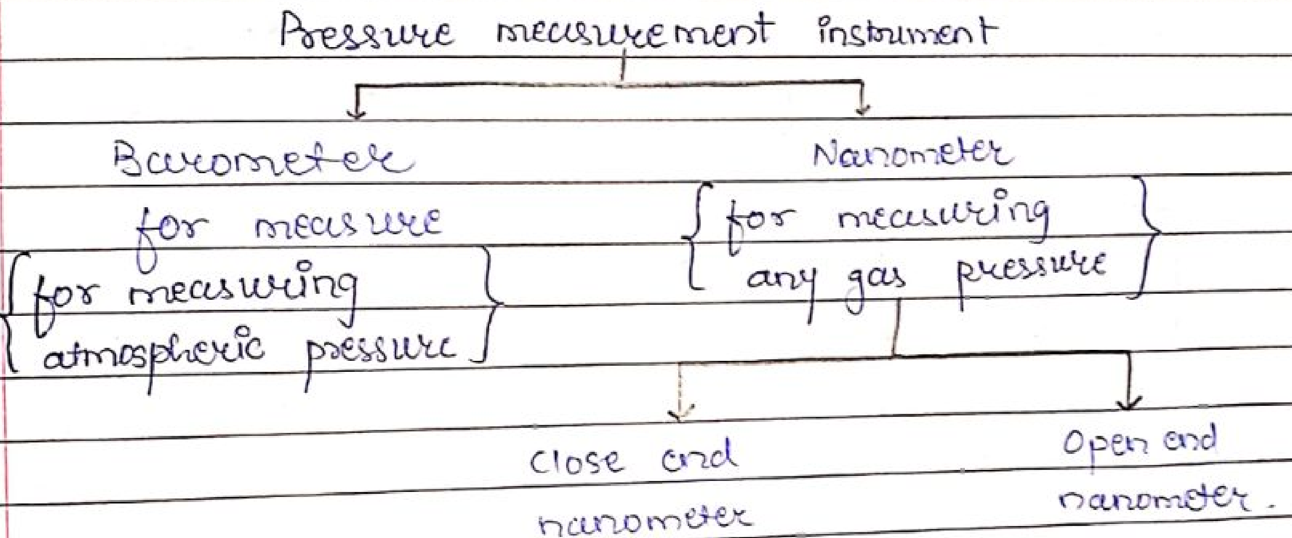
* unit of pressure =

1 atm = 1.013 × 10^5 Pascal = 760 mm Hg
= 76 cm Hg
= 760 torr

out of NCERT syllabus * measurement / calculation of pressure =



Gas laws :-

- the quantitative relationship between any two of variables (volume, mole, pressure and temperature) when other two are constant.

⟹ 1.) Boyle's law:

- at constant temperature for 'n' mole of gas pressure is inversely proportional to the volume of gas.

$$p \propto \frac{1}{V}$$ (at constant n and T)

$p = K \cdot \frac{1}{V}$ [here, K is proportional constant]

or $\boxed{pV = K}$ or $\boxed{p_1V_1 = p_2V_2}$ $pV = nRT = K$

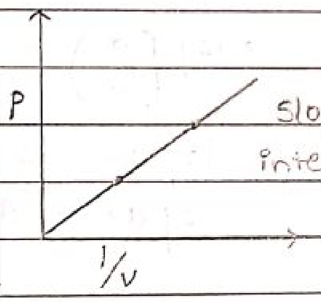
NOTE: here, isothermal curve should be obtained. (as temperature is constant.)

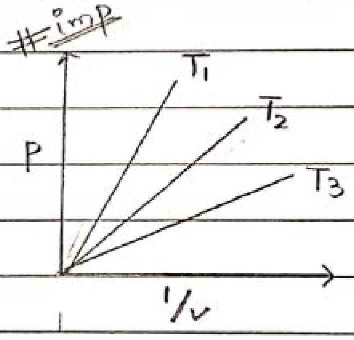
* Graphs for Boyle's law:

(a)

$p = K \cdot \frac{1}{V}$

$y = m x + C$



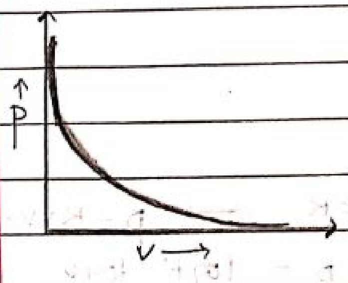


$y = p$ $x = 1/V$
$m = K$ $c = 0$

here, $T_1 > T_2 > T_3$
Reason: $pV = K$ or $pV = nRT = K$
or T↑ K↑ (Slope)

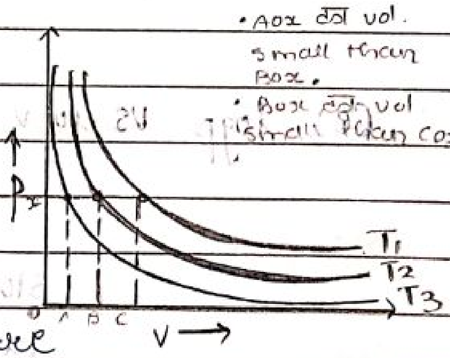
$$\underset{\underset{\text{axis}}{\downarrow}}{y} = \overset{\overset{\text{slope} = \tan\theta}{\uparrow}}{m}\underset{\underset{\text{axis}}{\downarrow}}{x} + C \rightarrow \text{intercept.}$$

(b)



p↑ V↓ $(P \propto 1/V)$

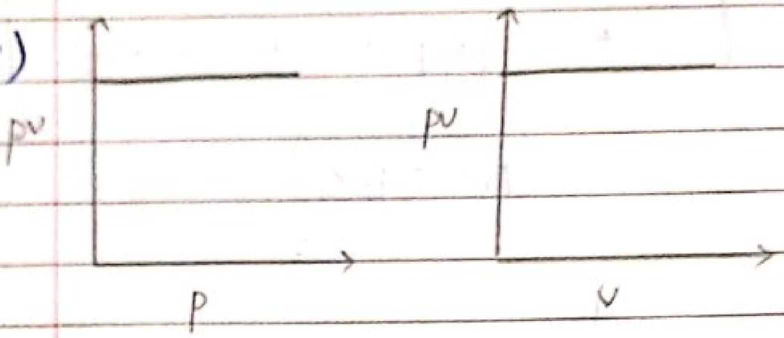
here, $T_1 > T_2 > T_3$



• Aox का vol. small than Box.
• Box का vol small than Cox.

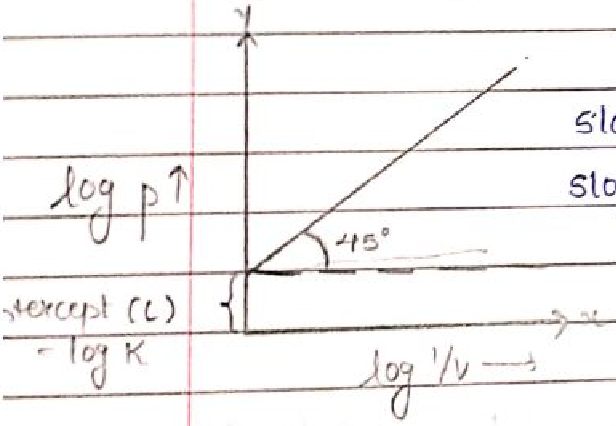
Reason: constant pressure volume↑ so, Slope or K↑.

(c)



here,
pv is constant.
hence;
$\underline{pv\uparrow\ p\downarrow}$
$pv\downarrow\ p\uparrow$.
also, $\underline{pv\uparrow\ v\downarrow}$

(d) log p vs log 1/v.



slope = 1
slope = tan45° = 1.

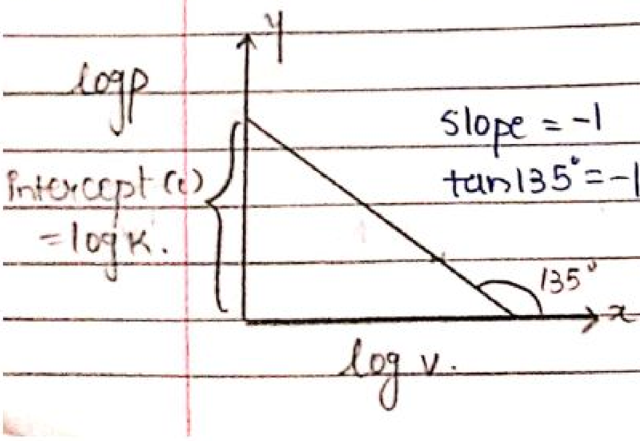
NOTE:
- $\log(a\times b) = \log a + \log b$
- $\log\left(\frac{a}{b}\right) = \log a - \log b$
- $\log a^2 = 2\log a$
- $\log a^3 = 3\log a$.

here; $pv = K \longrightarrow p = K\times\frac{1}{v}$

$$\log p = \log K + \log\frac{1}{v}$$

$$(y) = (c) + (mx)$$

(e) log p vs log v

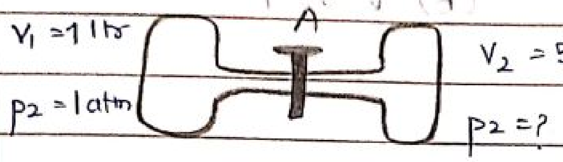


slope = −1
tan135° = −1

$pv = K \longrightarrow p = K\times v$.
$\log p = \log K - \log v$

or $\log p = -\log v + \log K$
$(y) = (-mx) + c$

Q:3 find out final pressure after opening of valve A in the following figure :

⟹ 

$P_1V_1 = P_2V_2$

$V_2 = 1$ ltr $+ 500$ ml

(opening valve) $= \underline{1.5 \text{ ltrs}}$

∴, $p_1V_1 = p_2V_2$

$1 \times 1 = p_2 \times 1.5$

$P_2 = \frac{1}{1.5}$ $\quad \frac{1\cancel{0}}{\cancel{15}}$ $\frac{2}{3}$

Q:4 a ballon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bars. If at 1 bar pressure gas occupies 2.27 ltr volume, upto what volume balloon can be expanded?

⟹ $P_1V_1 = P_2V_2$

$P_1 = 1$ bar $\quad V_1 = 2.27$ ltr

$P_2 = 0.2$ bar $\quad V_2 = ?$

∴, $(1)(2.27) = (0.2) \times V_2$

$V_2 = \frac{1 \times 2.27}{0.2} \longrightarrow \frac{1 \times \cancel{10} \times \overset{113.5}{\cancel{227}}}{\cancel{2} \times 10\cancel{0}} \Rightarrow \underline{11.35 \text{ ltr.}}$

⟹ 11.) **Charles law:-**

- at constant pressure volume of 'n' mole of gas is directly proportional to absolute temperature.

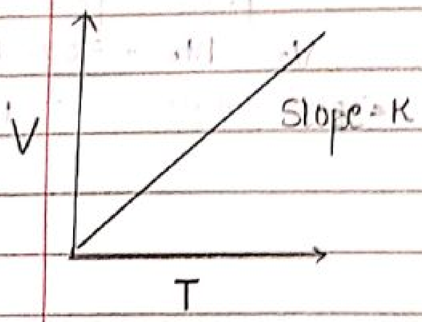
$\underline{V \propto T}$ (here, p and n are constant.)

$V = KT$ (here, K is a constant.)

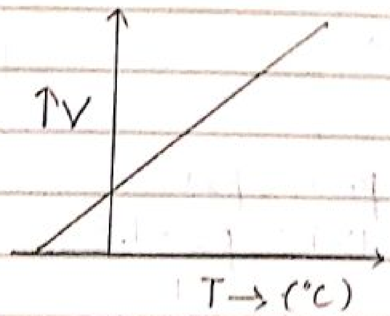
or $\boxed{\frac{V}{T} = K}$ or $\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$

* Different isobaric graph for Charle's law:

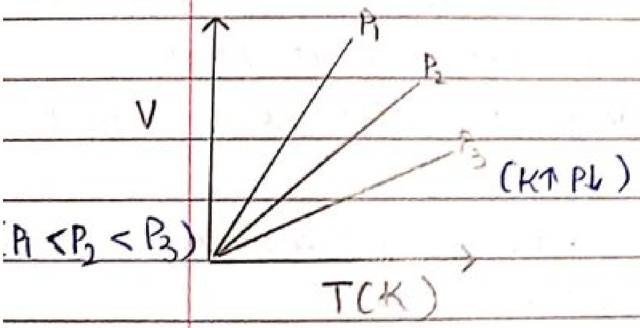
(a) V vs T



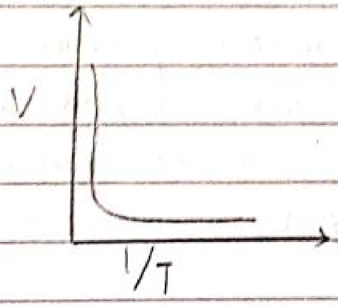
(b) V vs T



(c) V vs T



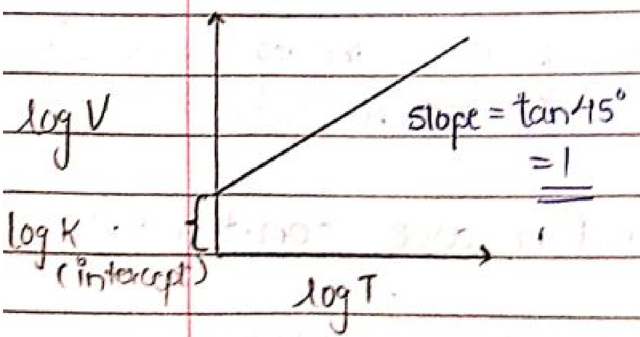
(d) V vs $1/T$



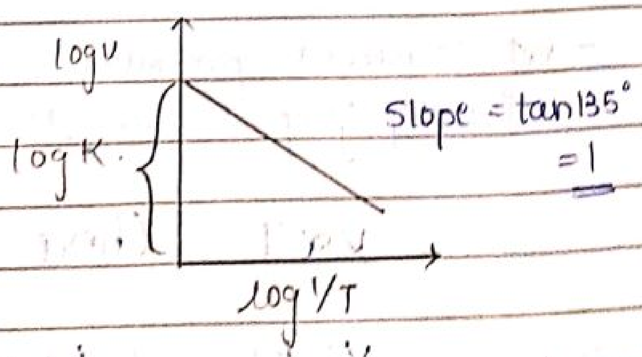
$\rightarrow PV = nRT$

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

(e) log V vs log T.



(f) log V vs log $1/T$



$V = KT$

$\underbrace{\log V}_{y} = \underbrace{\log K}_{c} + \underbrace{\log T}_{mx}$

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⇒ III.) Gay Lussac's law:

- at constant volume, pressure of 'n' mole of gas directly proportional to absolute temperature of gas.

$p \propto T$ (here V and n are constant)

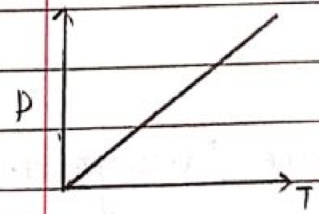
or $\boxed{p = KT}$ or $\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$

- at constant volume, graph will be isochoric.

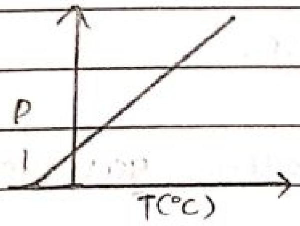
* Different isochoric graph for Gay Lussac law:

NOTE: all graphs are similar to Charles law graphs. (P instead V).

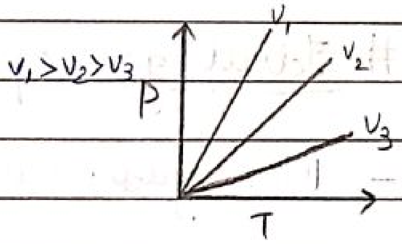
(a) P vs T



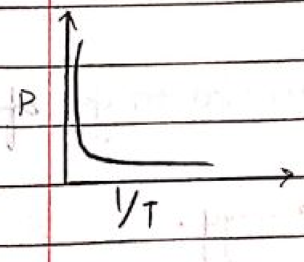
(b) P vs T



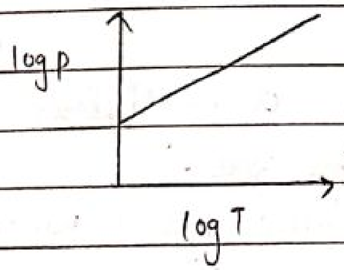
(c) P vs T.



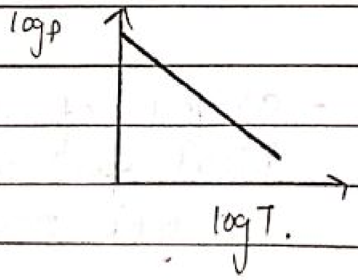
(d) P vs 1/T



(e) log P vs log T



(f) log p vs log 1/T.



⇛ IV.) Avogadro law:

- equal volume of any two gases at same temperature and pressure contains same number of molecules.

$V \propto n$ (here, p and T are constant)

or $\boxed{V = Kn}$ or $\boxed{\frac{V_1}{n_1} = \frac{V_2}{n_2}}$

Combined gas law:

- overall combination of Boyle's law, Charle's law, Gay lussac's law then obtained relation is known as combined gas law.

$$\frac{PV}{T} = K \longrightarrow \boxed{\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}}$$

Ideal gas equation:

- those gas which follow, gas law are ideal gases.

- in case of ideal gas, force of attraction between gaseous molecule is assumed zero.

- as force of attraction is negligible potential energy of ideal gas should be zero.

  ∴, total energy of ideal gas = kinetic energy.

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$$\boxed{PV = nRT.}$$

here; P = pressure of gas (atm) n = mole.
(container) V = volume of gas (ltr) R = universal gas constant.
T = absolute temperature (K)

* ideal gas equation is a relation between four variables and it describes the state of any gas. so this equation is also known as '<u>equation</u> of <u>state</u>'.

Different values for Universal Gas Constant:

- $R = \frac{PV}{nT} \Longrightarrow \frac{1\,atm \times 22.4\ ltr}{1\ mole \times 273\ K.} \Longrightarrow \boxed{0.0821\ \frac{atm\ ltr.}{mole\ K.}}$

- at 'STP' (273 K and 1 atm pressure):
  volume of 1 mole gas = 22.4 ltr.

- conversion to bar:
  1 bar = 1.01 atm
  $$\boxed{R = 0.0831\ \frac{bar\ ltr.}{mole\ K.}}$$

- conversion to energy:

$$\boxed{R = 8.31\ \frac{Joule}{mole\ K.}}$$

$$\boxed{R = 1.98 \simeq 2\ \frac{cal}{mole\ K.}}$$

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Q:5 Find volume of 32gm methane gas (assume ideal gas) at 27°C temperature and 380 mmHg pressure.

⟹ $CH_4 = 16gm$ 1 atm = 760 mmHg

$n = \frac{\cancel{32}}{\cancel{16}}$ $\underline{2 \text{ mole}}$ $= \frac{\cancel{380}}{\cancel{760}}$ $\frac{1}{2}$ or 0.5

K = 27 + 273 ⟶ 300K.

$$pV = nRT$$

$$\frac{1}{2} \cdot V = 2 \cdot 300 \cdot 0.0821$$

$$V = 1200 \cdot 0.0821$$ ∴ $\boxed{V = 98.52 \text{ ltr}}$

Q:6 A cylinder contain acetyline gas (C_2H_2) ($CH \equiv CH$) at 27°C and 4.05 megapascal (MPa) pressure. The pressure in cylinder after half the mass of gas used and now temperature has fallen to 12°C. Find final pressure of acetyline gas in MPa.

⟹ megapascal (MPa) = 10^6 vol. constant as fixed container

$P_1 \cancel{V_1} = n_1 \cancel{R_1} T_1$

$P_2 \cancel{V_2} = n_2 \cancel{R_2} T_2$

$\left[\begin{array}{l} R_1 = R_2 \\ V_1 = V_2 \end{array}\right]$

K = 27 + 273
= $\underline{300K}$.

K = 273 + 12
= $\underline{285K}$.

$$\therefore \frac{P_1}{n_1 T_1} = \frac{P_2}{n_2 T_2}$$

$$\frac{4.05}{n_1 \times 300} = \frac{x}{0.5 \times 285}$$

$$x = \frac{\overset{135}{\cancel{405}} \times 05 \times 285}{100 \times 10 \times \underset{100}{\cancel{300}}} \longrightarrow \frac{192375}{100000}$$

$\boxed{= 1.92}$ MPa.

135
5
675

Q:7 mod. Q:11 Pg:121

Gay lussac $\Rightarrow$ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$)

Q:8 → An unknown gas mass is 12gms (mol. wt: 120gm). Kept in closed vessel at 1atm pressure and 't°c' temperature.
→ If we increase temperature by 10°C than pressure rises by 10%. Find initial temperature of gas.

$\Longrightarrow$ $n = \frac{12}{120} = 0.1 \text{ mol}$ $P_i = 1 \text{ atm}$ $P_f = \frac{1\times 110}{100}$

$= 1.\cancel{0}1 \text{ atm}$

~~100 °C~~

$(P_1)(\cancel{V}) = \cancel{n}(\cancel{R})(t)$
$(P_2)(\cancel{V}) = \cancel{n}(R)(t)$

$\therefore, \frac{P_1}{T_1} = \frac{P_2}{T_2}$

$\frac{1}{t} = \frac{1.1}{t+10}$ $\longrightarrow$, $t+10 = (1.1)t$
$1.1t - t = 10$
$0.1t = 10$
$t = 10\times10$
$\boxed{t_1 = 100°C}$