CHAPTER 5.0

STATES OF MATTER

CHEMISTRY 1
SK015

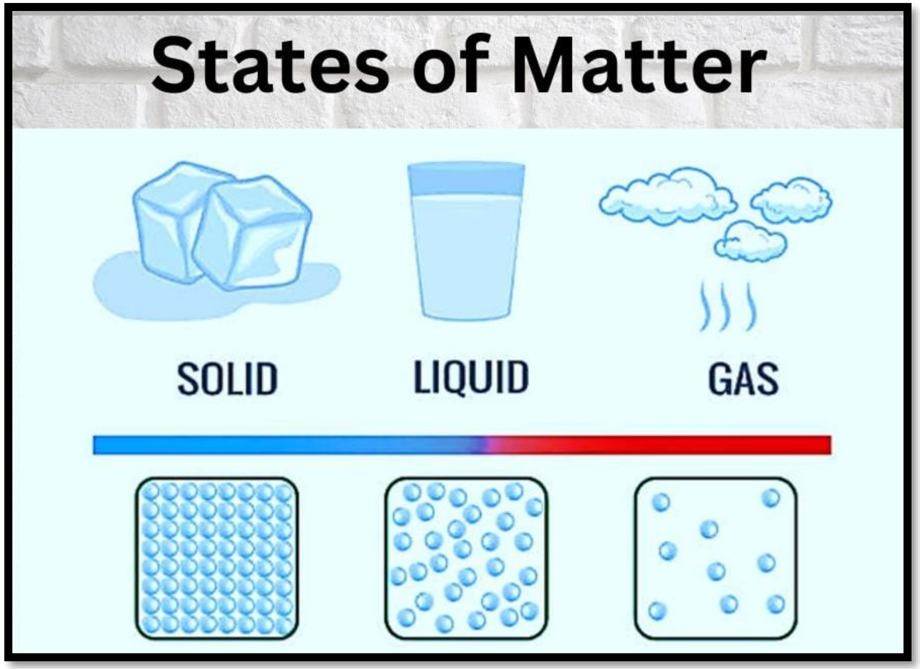
SESSION 2025/2026

STUDENT LEARNING TIME (SLT): LECTURE	
NON FACE-TO-FACE (PREPARATION)	FACE-TO-FACE (DURING CLASS)
3 HOURS	3 HOURS

CHEMISTRY UNIT, KMJ

5.0: STATES OF MATTER

- **5.1: GAS**
- **5.2: LIQUIDS**
- **5.3: SOLIDS**
- **5.4: PHASE DIAGRAM**



5.1 GAS

Teaching and learning outcomes

At the end of the lesson, student should be able to

5.1	Gas
a)	Explain qualitatively the basic assumptions of the kinetic molecular theory of gases for an ideal gas. (C2, C3)
b)	Define gas laws: (C1) i. Boyle's Law ii. Charles' Law iii. Avogadro's Law
c)	Interpret the graphs of Boyle' law and Charles' law. (C2, C3) *include sketch of graph
d)	Perform calculations involving gas laws and ideal gas equation. (C3)
e)	Determine molar mass of a volatile liquid using ideal gas equation. (C3) (Experiment 4: Charles' Law and the ideal gas law)

5.1 GAS

Teaching and learning outcomes

At the end of the lesson, student should be able to

5.1	**Gas**
f)	Define Dalton's law. (C1)
g)	Perform calculations using Dalton's Law. (C3)
h)	Explain qualitatively the ideal and non-ideal behaviours of gases in terms of intermolecular forces and molecular volume. **use van der Waals equation to explain the non-ideal behaviour* (C2, C3)
i)	Explain the conditions at which real gases approach the ideal behaviour. (C2, C3)
j)	Explain qualitatively the ideal and non-ideal behaviours of gases in terms of intermolecular forces and molecular volume. **use van der Waals equation to explain the non-ideal behaviour* (C2, C3)

KINETIC MOLECULAR THEORY OF GASES

Basic assumptions (Postulates) :

① Gas particles are in constant and random motion

- they frequently collide with one another and with the walls of container

② The volume of gas particles is negligible

- this is due to the size of gas particles are much smaller than the volume of container

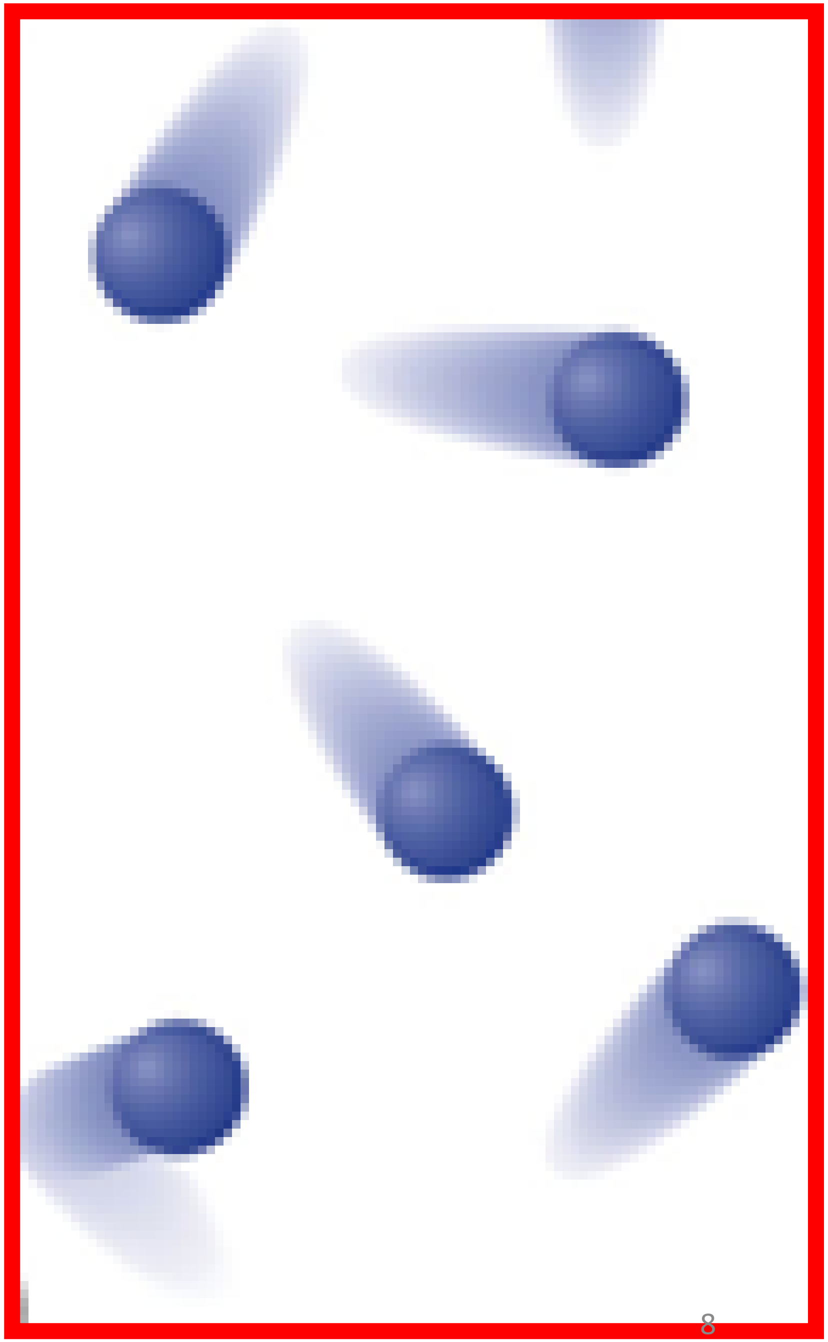
③ Intermolecular forces (attractive and repulsive force) between gas particles are negligible

④ Collisions between gas particles are elastic

- kinetic energies of particles before the the collisions are same as that of after collisions
- total kinetic energy is constant

⑤ The average kinetic energy of the gas particles is directly proportional to the absolute temperature

Gas pressure is caused by gas molecules colliding container's walls



THE GAS'S LAW

The physical behavior of gas can be described by four variables:

- Pressure (*P*)
- Volume (*V*)
- Temperature (*T*)
- Amount (number of moles) (*n*)

- Boyle's law
- Charles's law
- Avogadro's law

BOYLE’S LAW

At constant temperature, the volume occupied by a fixed amount of gas is inversely proportional to the pressure of a gas.

$V \alpha 1/P$ (T and n fixed)

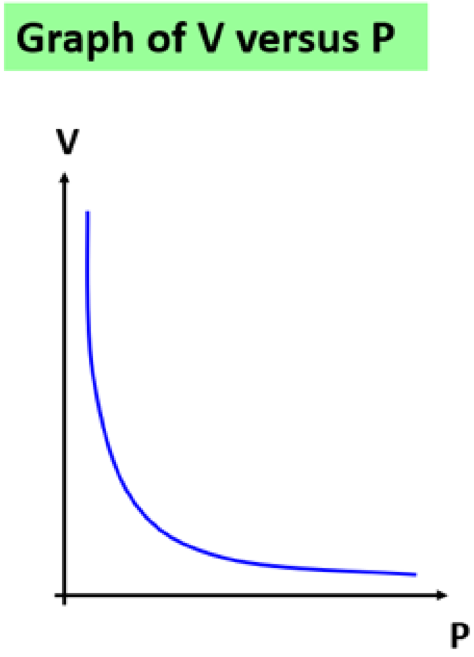
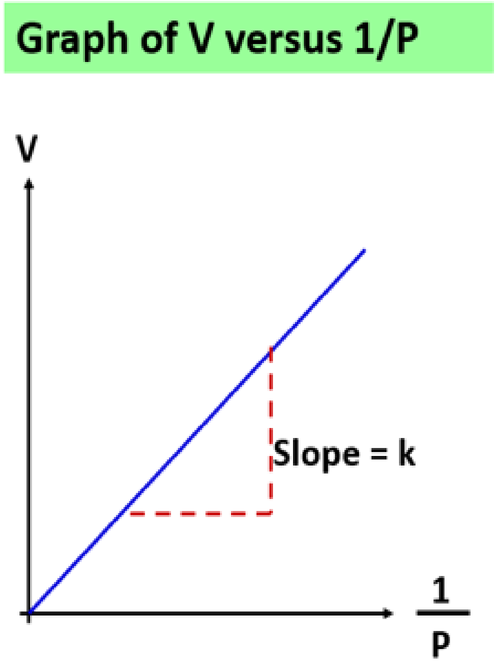
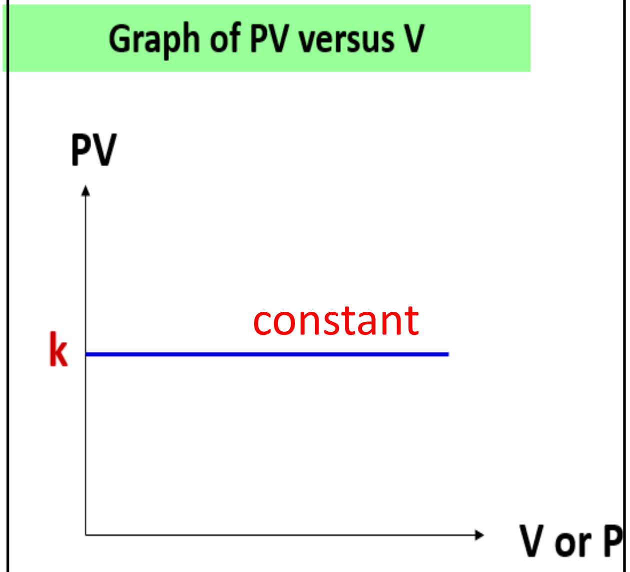
$PV = k$ Where k = constant

$$P_1V_1 = P_2V_2$$

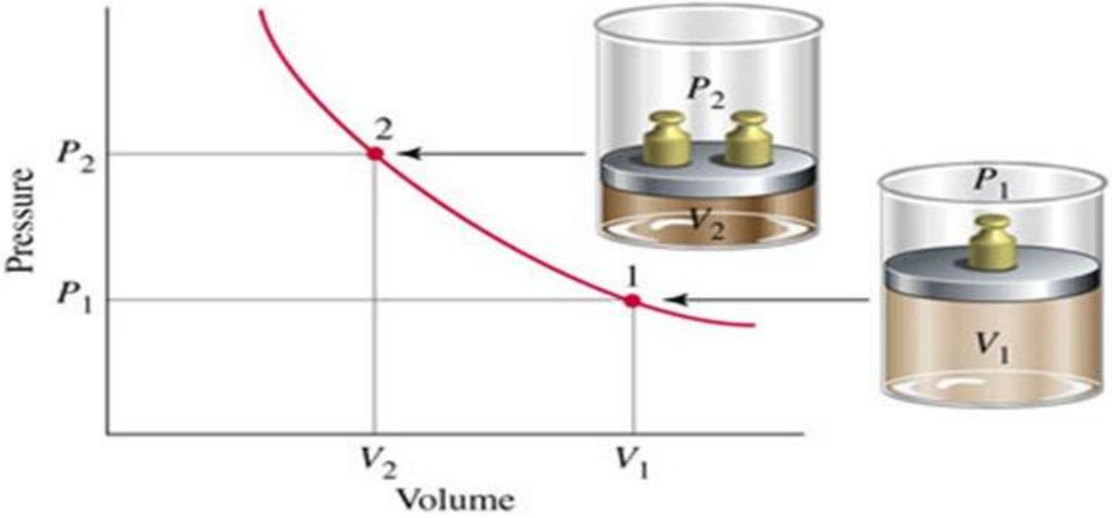
P_1 and V_1 = initial pressure and volume of the gas

P_2 and V_2 = final pressure and volume of the gas

Boyle's law can be represented by the following graphs:

 Graph of V versus P	 Graph of V versus 1/P	 Graph of PV versus V
Volume occupied by the gas is inversely proportional to the gas pressure. (As the gas pressure increases, volume of gas molecules decreases).	Volume occupied by the gas is directly proportional to 1/P. Therefore, the graph V against 1/P is a linear graph (passing through the origin).	Since PV is a constant, PV is independent of P or V.

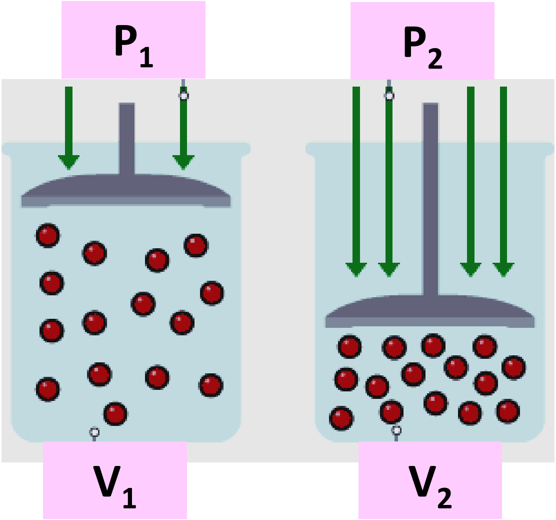
When gas is kept at constant temperature its pressure is inversely proportional to the volume.



A doubling of the pressure causes the volume of the gas to decrease to one-half its original value.

HOW THE KINETIC MOLECULAR THEORY EXPLAINS THE BOYLE'S LAW?

- **Gas pressure is a measure of the number of collisions between gas particles and the walls of container.**
- **The smaller the volume at constant *n* and *T*, the more crowded together the particles are and the greater the collisions.**
- **Thus, pressure of the gas increases as the volume decreases.**



EXAMPLE – 01

5.1

A sample of gas occupies 12.0 L under a pressure of 1.2 atm. What would its volume be if the pressure were increased to 2.4 atm?

EXAMPLE – 01

5.1

ANSWER

P_1 = 1.2 atm

P_2 = 2.4 atm

V_1 = 12.0 L

V_2 = ?

Using Boyle's law:

$P_1 \times V_1 = P_2 \times V_2$ $\qquad$ n , T constant

$$V_2 = \frac{P_1 \times V_1}{P_2} = \frac{1.2\ \cancel{\text{atm}} \times 12.0\ \text{L}}{2.4\ \cancel{\text{atm}}}$$

$$= \boxed{6.0\ \text{L}}$$

EXERCISE – 01

5.1

A cylinder contains a gas at 5.25 atm pressure. When the gas allowed to expand to a final volume of 12.5 L, the pressure drop to 1.85 atm. What was the original volume of the gas?

EXERCISE – 01

5.1

ANSWER

P_1 = 5.25 atm

P_2 = 1.85 atm

V_1 = ?

V_2 = 12.5 L

Using Boyle's law:

$P_1 \times V_1 = P_2 \times V_2$ n, T constant

$$V_1 = \frac{P_2 \times V_2}{P_1} = \frac{1.85\ \cancel{\text{atm}} \times 12.5\ \text{L}}{5.25\ \cancel{\text{atm}}}$$

$$= 4.40\ \text{L}$$

CHARLES'S LAW

At constant pressure, the volume occupied by a fixed amount of gas is directly proportional to its absolute temperature.

$V \propto T$ (P and n fixed)

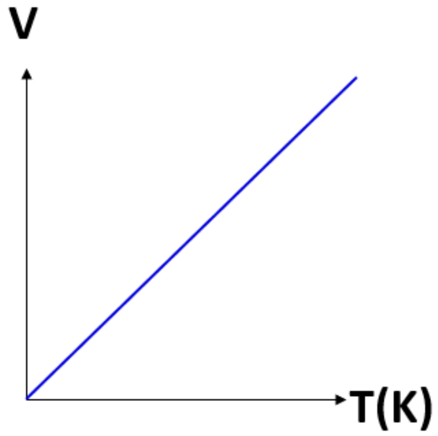
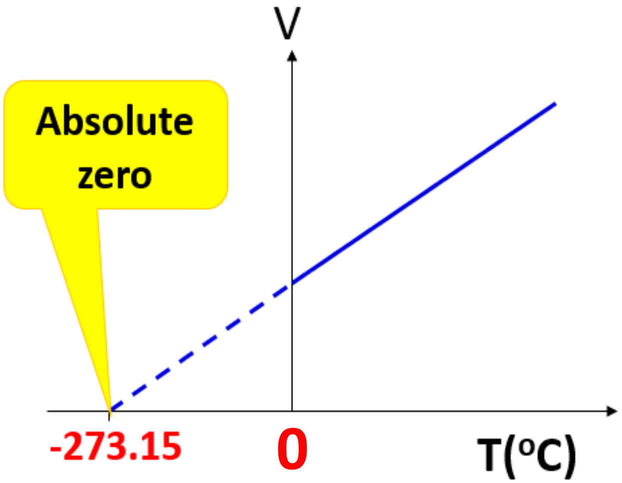
$$\frac{V}{T} = \text{k (constant)}$$

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

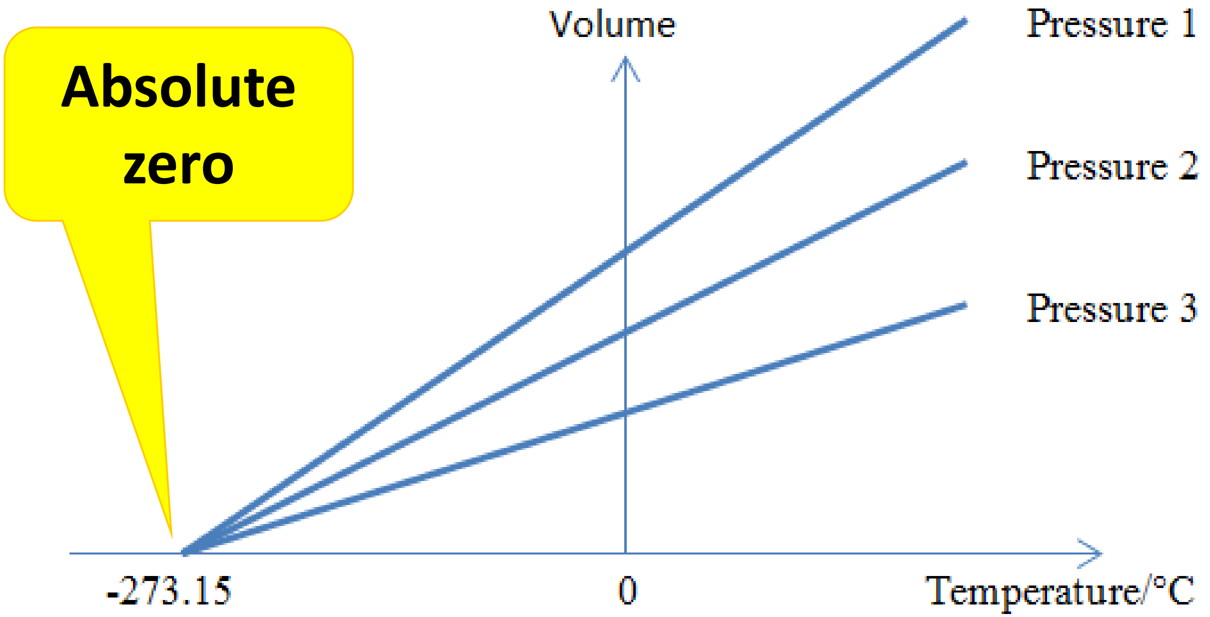
T in Kelvin (K)!

$T\text{ (K)} = t\ (^{o}\text{C}) + 273.15$

Charles’s law can be represented by the following graphs:

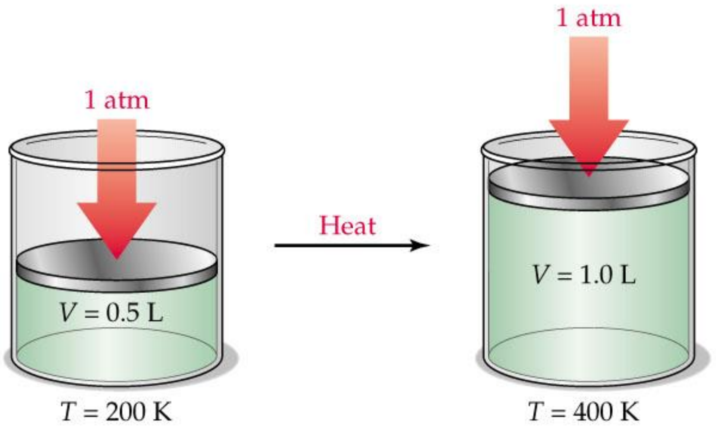
Graph of V versus T(K)  	Graph of V versus T(ºC) 
The volume occupied by the gas molecules is directly proportional to the absolute temperature (in Kelvin).	• As the temperature approaches zero, the volume occupied by the gas molecules approaches zero. • On extrapolation, the graph meets the temperature at −273.15°C. This temperature is called absolute zero.

Graph V against T (°C) at constant *P* and *n*



HOW THE KINETIC MOLECULAR THEORY EXPLAINS THE CHARLES'S LAW?

- **Temperature is a measure of the average kinetic energy of the gas particles.**
- **The higher the temperature at constant *n* and *P*, the faster the particles move and more space they need to move around in to avoid increasing their collisions with the walls of the container.**
- **Thus, volume increases as temperature increases.**



EXAMPLE – 02

5.1

A sample of carbon monoxide gas occupies 3.20 L at 125 ^{o}C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

EXAMPLE – 02

5.1

ANSWER

$V_1 = 3.20\ \text{L}$

$V_2 = 1.54\ \text{L}$

$T_1 = (125 + 273.15)\ \text{K}$

$= 398.15\ \text{K}$

$T_2 = ?$

Using Charles's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad n\ ,\ p\ \text{ fixed}$$

$$T_2 = \frac{V_2 \text{ x } T_1}{V_1} = \frac{1.54\ \cancel{\text{L}} \text{ x } 398.25\ \text{K}}{3.20\ \cancel{\text{L}}}$$

$$= \boxed{192\ \text{K}}$$

EXERCISE – 02

An engineer pumps air at 0°C into a newly designed piston–cylinder assembly. The volume measures 6.83 cm^3. At what temperature (in K) would the volume be 9.75 cm^3?

EXERCISE – 02

ANSWER

V_1 = 6.83 cm^3

V_2 = 9.75 cm^3

T_1 = (0 + 273.15) K

= 273.15 K

T_2 = ?

Using Charles's law:

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

n , *p* fixed

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{9.75 \cancel{\text{cm}^3} \times 273.15 \text{ K}}{6.83 \cancel{\text{cm}^3}}$$

= **390 K**

Chemistry unit

AVOGADRO'S LAW

- At constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of the gas present

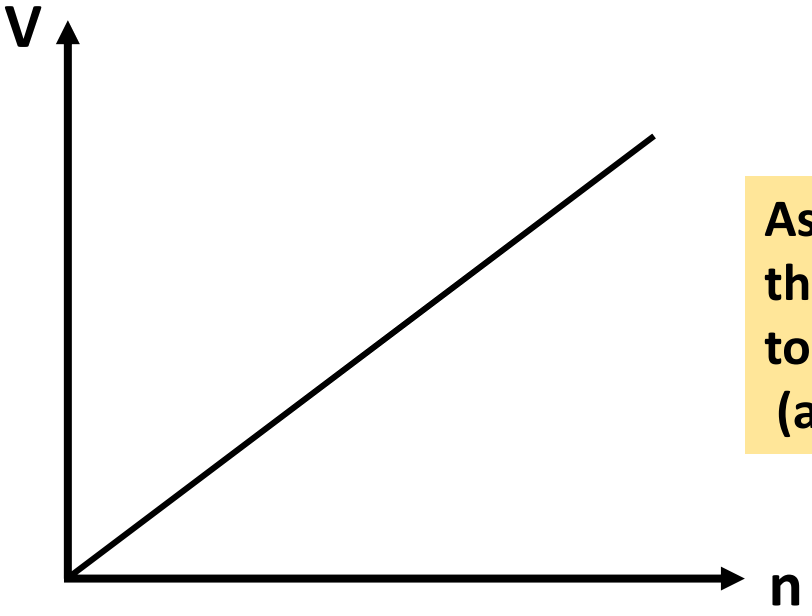
$V \alpha n$ (P and T fixed)

$$\frac{V}{n} = \text{k (constant)}$$

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

The graphical representation of Avogadro’s Law is shown below:

Graph Volume (V) against number of mole (n)



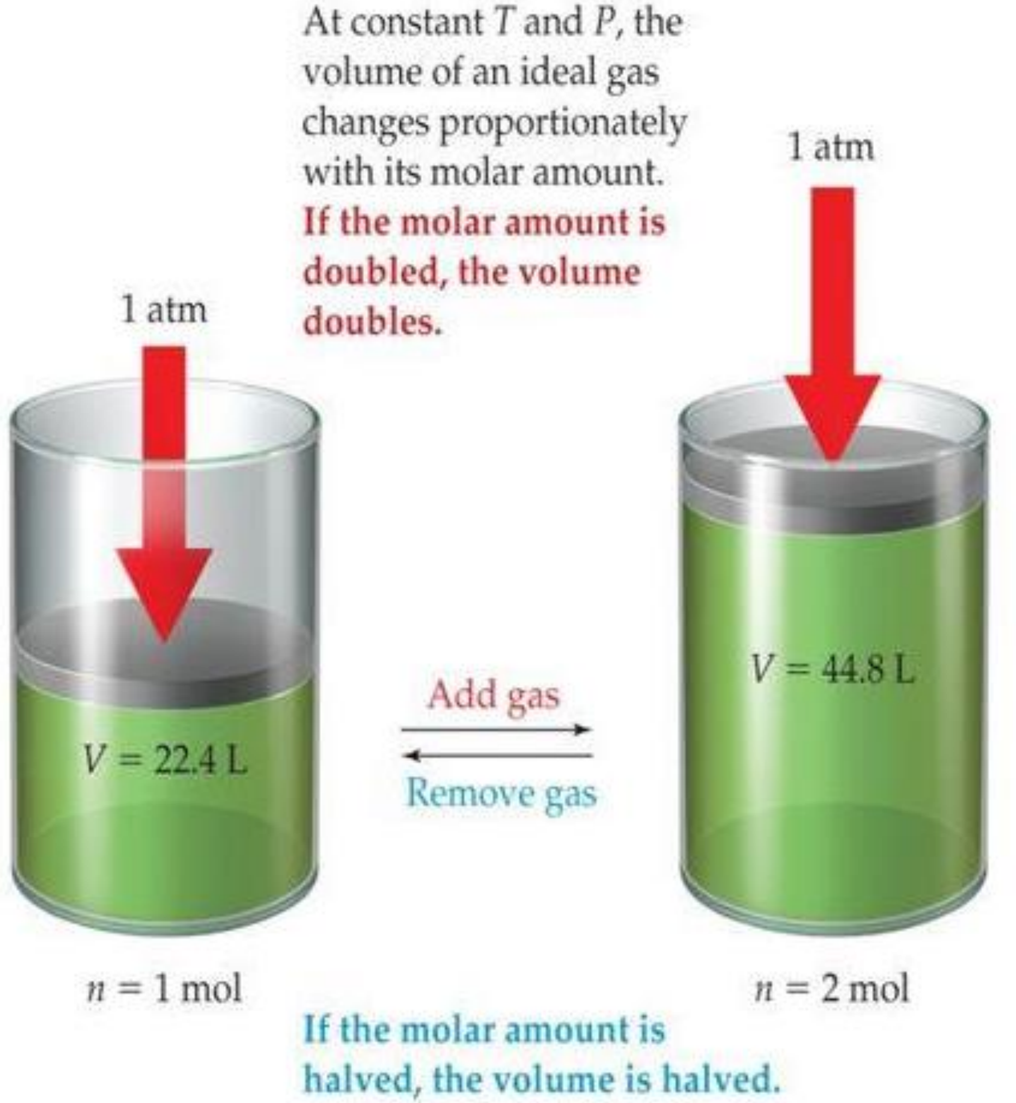
As we can observe from the graph,
the volume of a gas is directly proportional
to the number of moles of gas present
(at constant P and T).

The graph is straight line passing to the origin with positive slope.

HOW THE KINETIC MOLECULAR THEORY EXPLAINS THE AVOGADRO’S LAW?

- The more particles there are in the gas sample, the more volume the particles need at constant *P* and *T* to avoid increasing their collisions with the walls of the container.

- Thus, volume increases as amount (number of moles) of gas particles increases.



At constant T and P, the volume of an ideal gas changes proportionately with its molar amount. **If the molar amount is doubled, the volume doubles.**

If the molar amount is halved, the volume is halved.

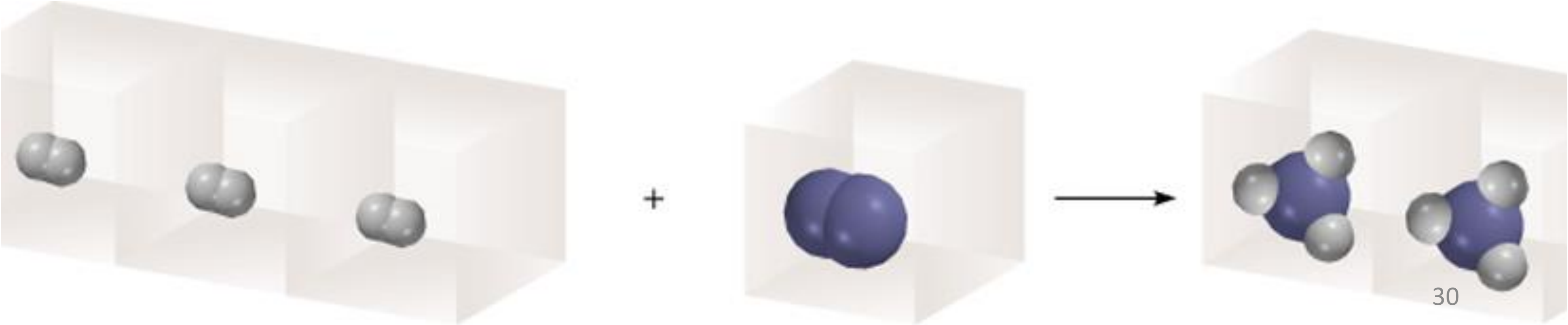
EXAMPLE – 03

5.1

Nitrogen and oxygen gases react to form ammonia gas in the reaction

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

If 500 mL N_2 at 1 atm and 25°C were available for reaction, what volume (in L) of H_2, at the temperature and pressure, would required in the reaction?



EXAMPLE – 03

5.1

According to Avogadro's Law:

$N_2(g)$ +	$3H_2(g)$ →	$2NH_3(g)$
1 mol	**3 mol**	**2 mol**
1 L	**3 L**	**2 L**
500 mL	**500 x 3 mL**	**500 x 2 mL**
	= 1500 L	**= 3000 L**
	= 1.5 L	

So, the volume of H_2 required = 1.5 L

EXERCISE – 03

5.1

Nitrogen monoxide (NO) combines with oxygen to form nitrogen dioxide.

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

If 1.0 L oxygen gas at 30.25°C and 0.975 atm is used and there is an excess of NO, what volume (in L) of NO gas at the same temperature and pressure will be converted to NO_2?

EXERCISE – 03

5.1

According to Avogadro's Law:

$2NO(g)$	+	$O_2(g)$	→	$2NO_2(g)$
2 mol		1 mol		2 mol
2 L		1 L		2 L

So, the volume of NO converted = 2.0 L

IDEAL GAS

- Hypothetical gas whose pressure–volume–temperature relationship can be accounted by the ideal gas equation.

- Any gas that obeys the ideal gas equation and has the properties as outlined by the Kinetic Molecular Theory

- Also called perfect gas

IDEAL GAS EQUATION

$$PV = nRT$$

Boyle's Law : $V \propto \frac{1}{P}$

Charles's Law : $V \propto T$

Avogadro's Law : $V \propto n$

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

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

$$PV = nRT$$

R = gas constant

$= 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$

$= 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

UNIVERSAL GAS CONSTANT (R)

$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})(22.414\ \text{L})}{(1\ \text{mol})(273.15\ \text{K})}$$

$$R = 0.08206\ \text{L}\ \ \text{atm} / (\text{mol}\ \ \text{K})$$

$$\downarrow$$

$$\text{L}\ \ \text{atm}\ \ \text{mol}^{-1}\ \ \text{K}^{-1}$$

EXAMPLE – 04

5.1

What pressure, in atm, is exerted by 54.0 grams of Xe in a 1.00–liter flask at 20^{o}C

EXAMPLE – 04

ANSWER

$$\text{Mole of Xe} = \frac{\text{Mass of Xe}}{\text{Molar mass of Xe}}$$

$$= \frac{54.0\ \text{g}}{131.29\ \text{g/mol}}$$

$$= 0.4113\ \text{mol Xe}$$

EXAMPLE – 04

5.1

By using ideal gas equation:

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

$$= \frac{0.4113\ \cancel{\text{mol}} \times 0.08206\ \text{atm}\bullet\text{L}\cancel{\text{mol}}^{-1}\cancel{\text{K}^{-1}} \times 293.15\ \cancel{\text{K}}}{1.00\ \cancel{\text{L}}}$$

$$= \boxed{9.89\ \text{atm}}$$

EXERCISE – 04

5.1

What is the volume of a gas balloon filled with 4.00 moles of He when the atmospheric pressure is 748 torr and the temperature us 30°C?

EXERCISE – 04

5.1

$$\text{Pressure } (P) = 748\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}}$$

$$= 0.984\ \text{atm}$$

$$\text{Temperature } (T) = (30 + 273.15)\ \text{K}$$

$$= 303.15\ \text{K}$$

EXERCISE – 04

5.1

By using ideal gas equation:

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$= \frac{4.00\ \cancel{\text{mol}} \times 0.08206\ \cancel{\text{atm}}\bullet\text{L}\cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}} \times 303.15\ \cancel{\text{K}}}{0.984\ \cancel{\text{atm}}}$$

$$= 101.1\ \text{L}$$

FURTHER APPLICATION OF IDEAL GAS EQUATION

☐ Calculation of:

- Molar mass of a gas
- Density of a gas

PV = nRT

Determining the Molar Mass of a Gas from the Ideal Gas Law

$$PV = nRT$$

$$PV = \frac{m}{\mathcal{M}}RT \qquad \left[n = \frac{m}{\mathcal{M}}\right]$$

$$\mathcal{M} = \frac{mRT}{PV}$$

$$\mathcal{M} = \frac{mRT}{PV} \quad \text{or} \quad \mathcal{M} = \frac{dRT}{P}$$

m : Mass of the gas in grams (g)

d : density of gas

$\mathcal{M}$: Molar mass of the gas (Unit: g mol^{-1})

Determining the Density of a Gas from the Ideal Gas Law

$$PV = nRT$$

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

$$\left[n = \frac{m}{\mathcal{M}} \right]$$

$$P = \frac{mRT}{V\mathcal{M}}$$

$$\left[d = \frac{m}{V} \right]$$

$$P = \frac{d\,RT}{\mathcal{M}}$$

$$d = \frac{P\mathcal{M}}{RT}$$

m : Mass of the gas in grams (g)

$\mathcal{M}$: Molar mass of the gas (Unit: g mol^{-1})

d : density of gas

EXAMPLE – 5

5.1

An organic chemist isolates from a petroleum sample a colorless liquid with properties of cyclohexane (C_6H_{12}). He obtains the following data to determine its molar mass ($\mathcal{M}$):

Volume (*V*) of sample gas = 213 mL

Mass sample gas(*m*) = 0.582 g

***T* = 100.0°C** **_P_ = 754 torr**

Calculate molar mass of cyclohexane under these Conditions.

EXAMPLE – 5

5.1

ANSWER

$V = 0.213\ \text{L}$

$m = 0.582\ \text{g}$

$$T = (100 + 273.15)\ \text{K}$$
$$= 373.15\ \text{K}$$

$$P = 754\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}}$$
$$= 0.992\ \text{atm}$$

By using the ideal gas equation:

$$PV = nRT$$

$$\mathcal{M} = \frac{mRT}{PV}$$

$$= \frac{0.582\ \text{g} \times 0.08206\ \cancel{\text{atm}}\ \cancel{\text{L}}\ \text{mol}^{-1}\cancel{\text{K}^{-1}} \times 373.15\ \cancel{\text{K}}}{0.992\ \cancel{\text{atm}} \times 0.213\ \cancel{\text{L}}}$$

$$= \boxed{84.3\ \text{g mol}^{-1}}$$

EXERCISE – 5

5.1

A 1.27 g sample of an oxide of nitrogen, believed to be either NO or N_2O, occupies a volume of 1.07 L at 25°C and 737 mmHg. Calculate molar mass of this oxide. Determine the molecular formula of this oxide.

EXERCISE –5

5.1

$V = 1.07\ \text{L}$

$m = 1.27\ \text{g}$

$T = (25 + 273.15)\ \text{K}$
$= 298.15\ \text{K}$

$$P = 737\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}}$$
$$= 0.970\ \text{atm}$$

By using the ideal gas equation:

$PV = nRT$

$$\mathcal{M} = \frac{mRT}{PV}$$

$$= \frac{1.27\ \text{g} \times 0.08206\ \cancel{\text{atm}}\ \text{L mol}^{-1}\cancel{\text{K}^{-1}} \times \cancel{2}98.15\ \cancel{\text{K}}}{0.970\ \cancel{\text{atm}} \times 1.07\ \cancel{\text{L}}}$$

$= 29.9\ \text{g mol}^{-1}$
(the oxide is NO)

EXAMPLE – 6

5.1

What is the density of oxygen gas (O_2) at 298 K and 0.987 atm?

EXAMPLE – 6

5.1

$$d = \frac{P\mathcal{M}}{RT}$$

$$= \frac{0.987\ \cancel{\text{atm}}\ \times\ 32.0\ \text{g}\cancel{\text{mol}^{-1}}}{0.08206\ \cancel{\text{atm}}\ \text{L}\ \cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}}\ \times\ 298\ \cancel{\text{K}}}$$

$$= \boxed{1.29\ \text{g L}^{-1}}$$

EXAMPLE – 7

The density of phosphorus vapor at 310^{o}C and 775 mmHg is 2.64 g/L. What is the molecular formula of the phosphorus under these conditions?

EXAMPLE – 7

5.1

ANSWER

$d = 2.64\ \text{gL}^{-1}$

$T = (310 + 273.15)\ \text{K}$

$\quad = 583.15\ \text{K}$

$$P = 775\ \cancel{\text{mmHg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mmHg}}}$$

$$= 1.0197\ \text{atm}$$

By using ideal gas law:

$$PV = nRT$$

$$\mathcal{M} = \frac{dRT}{P}$$

$$= \frac{2.64\ \text{g}\cancel{\text{L}^{-1}} \times 0.08206\ \cancel{\text{atm}}\ \cancel{\text{L}}\ \text{mol}^{-1}\cancel{\text{K}^{-1}} \times 583.15\ \cancel{\text{K}}}{1.0197\ \cancel{\text{atm}}}$$

$$= 124\ \text{g mol}^{-1}$$

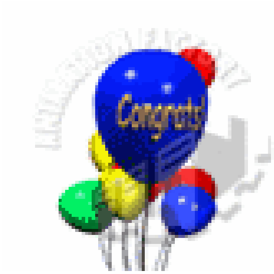
EXAMPLE – 7

5.1

ANSWER

$$\frac{\text{Molar mass of phoshprus molecule}}{\text{Molar mass of phosphorus element}} = n \text{ (whole number)}$$

$$\frac{124 \text{ g mol}^{-1}}{31.0 \text{ g mol}^{-1}} = 4$$



Molecular Formula = P_4

DALTON'S LAW OF PARTIAL PRESSURE

- In a mixture of unreacting gases, the total pressure is the sum of the partial pressure exerted by each gas in the mixture (at constant V and T)

$$P_{Total} = P_A + P_B + P_C + \ldots\ldots$$

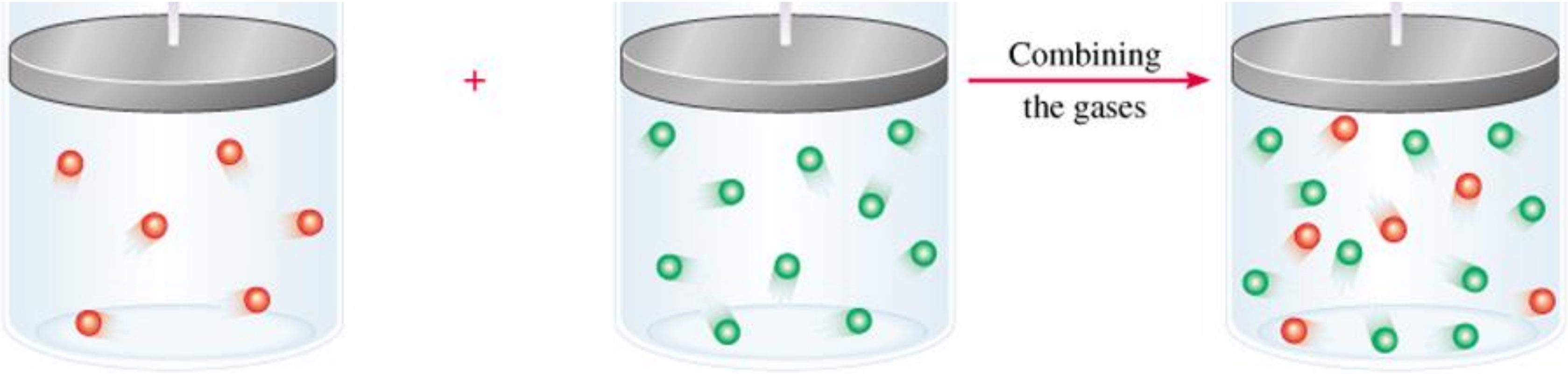
EXAMPLE:

$$P_{Total} = P_{N_2} + P_{He} + P_{Ar}$$

John Dalton (1766-1844).
English chemist, mathematician, and philosopher

PARTIAL PRESSURE

- Pressures exerted by individual gas components in the gas mixture



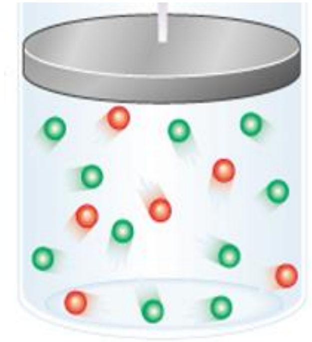
$P_{N_2} = 1$ atm

$P_{Ar} = 2$ atm

$P_{Total} = P_{N_2} + P_{Ar}$

$= 3$ atm

Consider a case in which two gases, A and B, are in a container of fixed volume V:



The pressure exerted by gas A and gas B is:

$$P_A = \frac{n_A RT}{V}$$

[n_A : moles of gas A]

$$P_B = \frac{n_B RT}{V}$$

[n_B : moles of gas B]

According to Dalton's law:

P_{Total} = Partial pressure of gas **A** + Partial pressure of gas **B**

$$P_{Total} = P_A + P_B$$

$$= \frac{n_A RT}{V} + \frac{n_B RT}{V}$$

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

$$P_{Total} = (n_{Total})\frac{RT}{V}$$

Partial pressure of gases can be expressed in terms of mole fraction:

$$\text{Mole fraction } (X_A) = \frac{\text{moles of gas A}}{\text{Total number of moles of all gases}}$$

$$X_A = \frac{\left[\dfrac{P_A \cancel{V}}{\cancel{RT}}\right]}{\left[\dfrac{P_{Total} \cancel{V}}{\cancel{RT}}\right]}$$

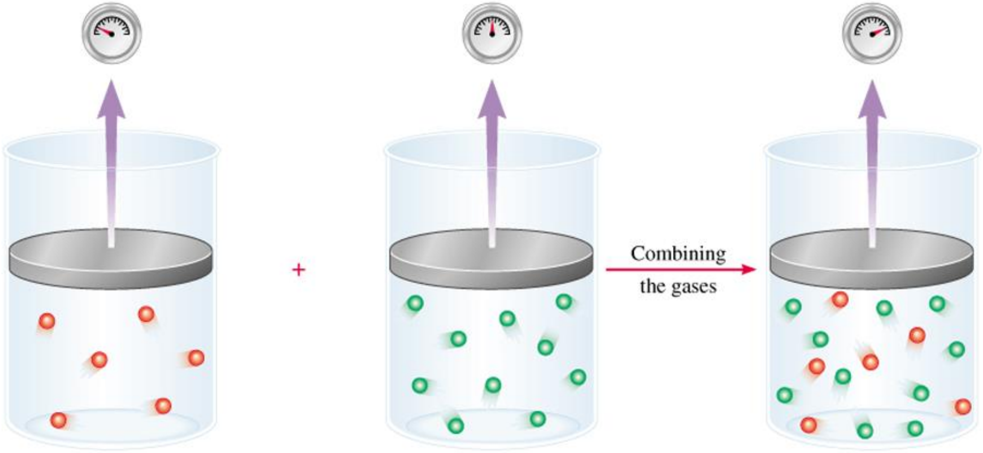
$$X_A = \frac{P_A}{P_{Total}}$$

$$\boxed{P_A = X_A P_{Total}}$$

EXAMPLE – 8

5.1

A gaseous mixture made from 6.00 g O_2 and 9.00 g CH_4 is placed in a 15.0 L vessel at 0°C. What is the partial pressure of each gas, and what is the total pressure in the vessel?



EXAMPLE – 8

Oxygen gas (O_2):

$$\text{Mol} = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1}\text{)}}$$

$$\text{Mole of } O_2 = \frac{6.00 \text{ g}}{32.00 \text{ g mol}^{-1}}$$

$$= 0.188 \text{ mol } O_2$$

EXAMPLE – 8

By using the ideal gas equation:

$$PV = nRT$$

$$P_{O_2} = \frac{n_{O_2}RT}{V}$$

$$= \frac{0.188\ \cancel{\text{mol}} \times 0.08206\ \text{atm}\ \cancel{\text{L}}\ \cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}} \times 273\ \cancel{\text{K}}}{15.0\ \cancel{\text{L}}}$$

$$= \boxed{0.281\ \text{atm}}$$

EXAMPLE – 8

5.1

Methane gas (CH_4):

$$\text{Mol} = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1}\text{)}}$$

$$\text{Mole of } CH_4 = \frac{9.00\text{ g}}{16.01\text{ g mol}^{-1}}$$

$$= 0.563\text{ mol } CH_4$$

EXAMPLE – 8

By using the ideal gas equation:

$$PV = nRT$$

$$P_{CH_4} = \frac{n_{CH_4}\, RT}{V}$$

$$= \frac{0.563\ \cancel{\text{mol}}\ \times\ 0.08206\ \text{atm}\ \cancel{\text{L}}\ \cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}}\ \times\ 273\ \cancel{\text{K}}}{15.0\ \cancel{\text{L}}}$$

$$= \boxed{0.841\ \text{atm}}$$

EXAMPLE – 8

By using the Dalton’s Law of partial pressure:

$$P_{Total} = P_{O_2} + P_{CH_4}$$

$$= 0.281 \text{ atm} + 0.841 \text{ atm}$$

$$= 1.122 \text{ atm}$$

EXERCISE – 6

What is the total pressure exerted by a mixture of 2.00 g of H_2 and 8.00 g of N_2 at 273 K in a 10.0 L vessel.

EXERCISE – 6

5.1

ANSWER

H_2 gas:

$$\text{Mol} = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1}\text{)}}$$

$$\text{Mole of } H_2 = \frac{2.00\ \text{g}}{2.02\ \text{g mol}^{-1}}$$

$$= 0.990\ \text{mol}\ H_2$$

EXERCISE – 6

5.1

H_2 gas:

By using the ideal gas equation:

$$PV = nRT$$

$$P_{H_2} = \frac{n_{H_2}\ RT}{V}$$

$$= \frac{0.990\ \text{m}\cancel{\text{o}}\text{l} \times 0.08206\ \text{atm}\ \cancel{\text{L}}\ \cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}} \times 273\ \cancel{\text{K}}}{10.0\ \cancel{\text{L}}}$$

$$= 2.218\ \text{atm}$$

EXERCISE – 6

5.1

N_2 gas:

$$\text{Mol} = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1}\text{)}}$$

$$\text{Mole of } N_2 = \frac{8.00\ \text{g}}{28.02\ \text{g mol}^{-1}}$$

$$= 0.286\ \text{mol } N_2$$

EXERCISE – 6

5.1

N_2 gas:

By using the ideal gas equation:

$$PV = nRT$$

$$P_{N_2} = \frac{n_{N_2} RT}{V}$$

$$= \frac{0.286\ \cancel{\text{mol}} \times 0.08206\ \text{atm}\ \cancel{\text{L}}\ \cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}} \times 273\ \cancel{\text{K}}}{10.0\ \cancel{\text{L}}}$$

$$= 0.641\ \text{atm}$$

EXERCISE – 6

$P_{H_2} = 2.218 \text{ atm}$

$P_{N_2} = 0.641 \text{ atm}$

By using the Dalton’s Law of partial pressure:

$$P_{Total} = P_{H_2} + P_{N_2}$$

$$= 2.218 \text{ atm} + 0.641 \text{ atm}$$

$$= 2.86 \text{ atm}$$

EXERCISE – 7

A 4.00 L flask containing He at 6.00 atm is connected to a 2.00 L flask containing N_2 at 3.00 atm and the gases are allowed to mix.

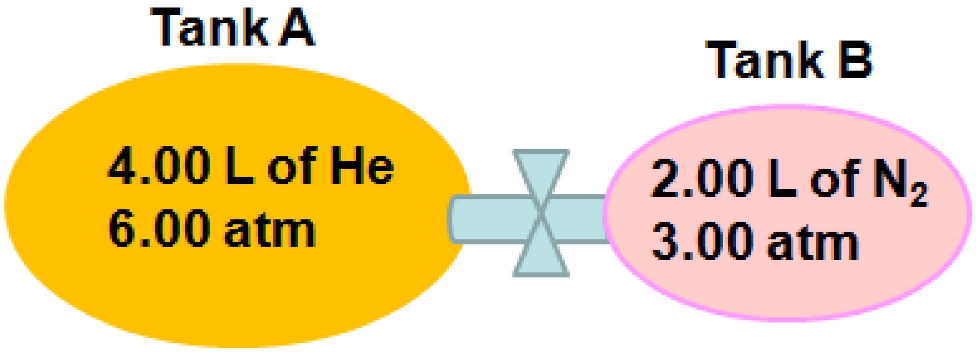
(a) Find the partial pressure of each gas after the gases are allowed to mix.

(b) Find the total pressure of the mixture.

(c) What is the mole fraction of helium?

EXERCISE – 7

(a)



Tank A

4.00 L of He
6.00 atm

Tank B

2.00 L of N_2
3.00 atm

Partial pressure of He:

Using Boyle's law:

$$P_1V_1 = P_2V_2 \qquad n\text{ , }T\ \textbf{constant}$$

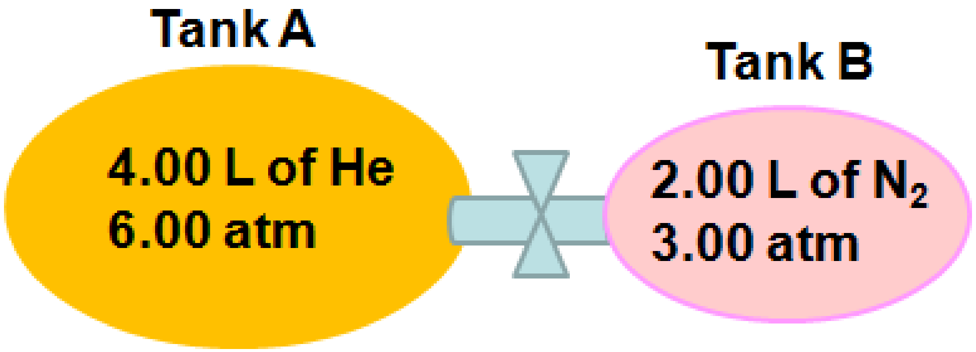
$$P_2 = \frac{P_1 \text{ x } V_1}{V_2} = \frac{6.00 \text{ atm x } 4.00\ \cancel{\text{L}}}{6.00\ \cancel{\text{L}}}$$

$$P_2 = P_{He} = = 4.00 \text{ atm}$$

EXERCISE – 7

5.1

(a)



Tank A

4.00 L of He
6.00 atm

Tank B

2.00 L of N_2
3.00 atm

Partial pressure of N_2:

Using Boyle's law:

$$P_1V_1 = P_2V_2 \qquad n\text{ , }T\text{ constant}$$

$$P_2 = \frac{P_1\,V_1}{V_2} = \frac{3.00\text{ atm } \times\ 2.00\ \cancel{\text{L}}}{6.00\ \cancel{\text{L}}}$$

$$P_2 = P_{N_2} = 1.00\text{ atm}$$

EXERCISE – 7

(b)

$P_{He} = 4.00 \text{ atm}$ $P_{N_2} = 1.00 \text{ atm}$

By using the Dalton's Law of partial pressure:

$$P_T = P_{He} + P_{N_2}$$

$$= 4.00 \text{ atm} + 1.00 \text{ atm}$$

$$= 5.00 \text{ atm}$$

EXERCISE – 7

5.1

(c)

P_{He} = 4.00 atm $\qquad$ P_T = 5.00 atm

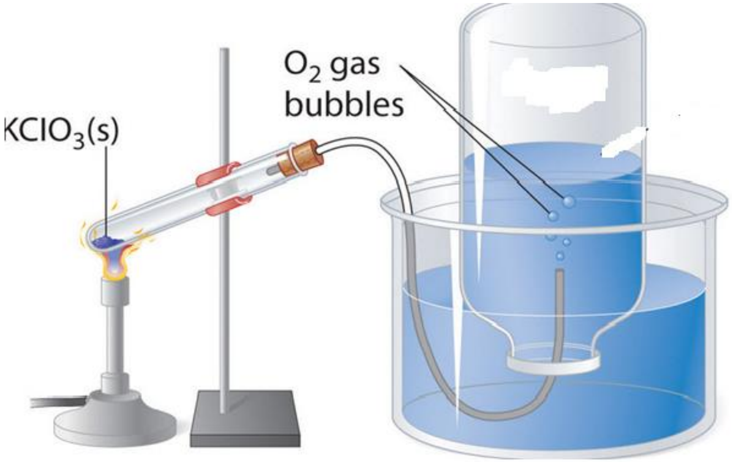
$$X_{He} = \frac{P_{O_2}}{P_T}$$

$$= \frac{4.00\ \cancel{\text{atm}}}{5.00\ \cancel{\text{atm}}}$$

$$= \ 0.800$$

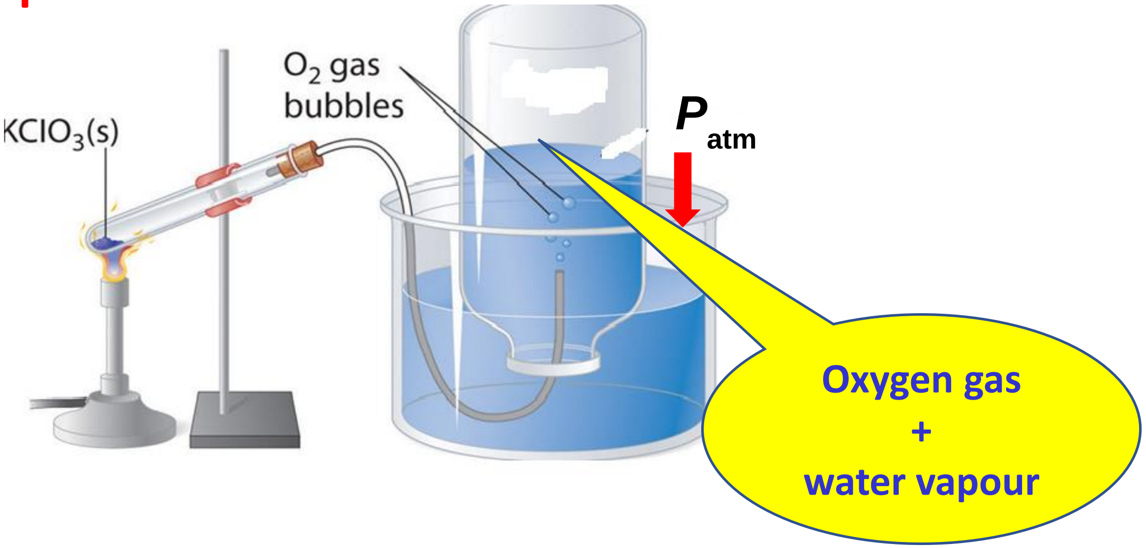
WATER DISPLACEMENT METHOD

APPLICATION OF DALTON'S LAW
IN COLLECTING A GAS OVER WATER



$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

Example:



Oxygen gas
+
water vapour

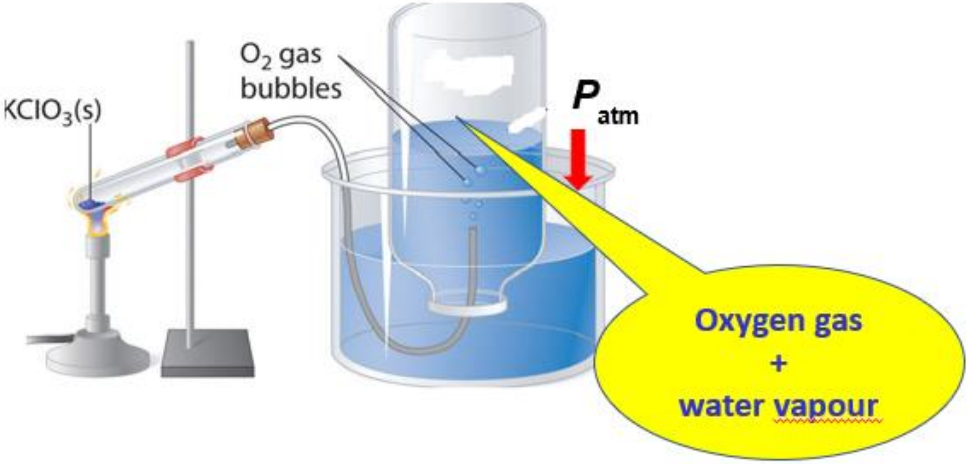
$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

- **When a gas is collected "over water", a mixture of the particular gas and water vapor is collected.**

- **By using Dalton's law:**

$$P_{Total} = P_{gas} + P_{H_2O}$$

Example:



$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

According to **Dalton's law of partial pressure**:

$$P_{Total} = P_{O_2} + P_{H_2O}$$

$$P_{O_2} = P_{Total} - P_{H_2O}$$

P_{H_2O} : pressure of water vapor

(at 25^{o}C = 23.8 mmHg = 23.8 torr)

Vapor Pressure of Water at Diffrerent Temperature

Temperature (°C)	Pressure (mmHg)
10	9.2
20	17.5
25	23.8
30	31.8
40	55.3
50	92.5
60	149.4
70	233.7
80	355.1

The vapour pressure of water depends on temperature, T. It increases with increasing T because higher T cause more water molecules to evaporate

EXAMPLE – 9

5.1

Consider the reaction below:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

A sample of 5.45 liters of oxygen is collected over water at a pressure of 735.5 torr and at a temperature of 25°C. How many grams of oxygen have been collected?

At 25°C, the vapor pressure of water = 23.8 torr

EXAMPLE – 9

5.1

According to Dalton's law of partial pressure:

ANSWER

$$P_T = P_{O_2} + P_{H_2O}$$

$$P_{O_2} = P_T - P_{H_2O}$$

$$= 735.5 \text{ torr} - 23.8 \text{ torr}$$

$$= 711.7 \text{ torr}$$

convert to atm

$$= 711.7 \text{ } \cancel{\text{torr}} \times \frac{1 \text{ atm}}{760 \text{ } \cancel{\text{torr}}}$$

$$= 0.936 \text{ atm}$$

EXAMPLE – 9

5.1

By using ideal gas equation:

$$PV = nRT$$

$$n_{O_2} = \frac{P_{O_2}V}{RT}$$

$$\text{Mole of } O_2 = \frac{0.936\ \cancel{\text{atm}} \times 5.45\ \cancel{\text{L}}}{0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\cancel{\text{K}^{-1}} \times 298.15\ \cancel{\text{K}}}$$

$$= 0.208\ \text{mol } O_2$$

EXAMPLE – 9

5.1

Mass of O_2 = mole of O_2 x molar mass of O_2

= 0.208 mol x 32.0 g/mol

= 6.66 g O_2

EXERCISE – 8

A nitrogen sample occupies 469 mL at STP. If the same sample were collected over water at 25°C and 750 torr, what would be the volume (in mL) of the gas sample?

(The vapor pressure of water at 25°C is 23.8 torr)

EXERCISE – 8

5.1

At STP, 1 22.4 L occupied by 1 mol N_2

Therefore,

$$0.469\ \cancel{\text{L}} \quad \text{x} \quad \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}}$$

$$= \quad 2.09 \text{ x } 10^{-2}\ \text{mol } N_2$$

EXERCISE – 8

According to Dalton's law of partial pressure:

$$P_T = P_{N_2} + P_{H_2O}$$

$$P_{N_2} = P_T - P_{H_2O}$$

$$= 750 \text{ torr} - 23.8 \text{ torr}$$

$$= 726.2 \text{ torr}$$

convert to atm

$$= 726.2 \text{ }\cancel{\text{torr}} \times \frac{1 \text{ atm}}{760 \text{ }\cancel{\text{torr}}}$$

$$= 0.956 \text{ atm}$$

EXERCISE – 8

5.1

By using ideal gas equation:

$$PV = n_{N_2}RT$$

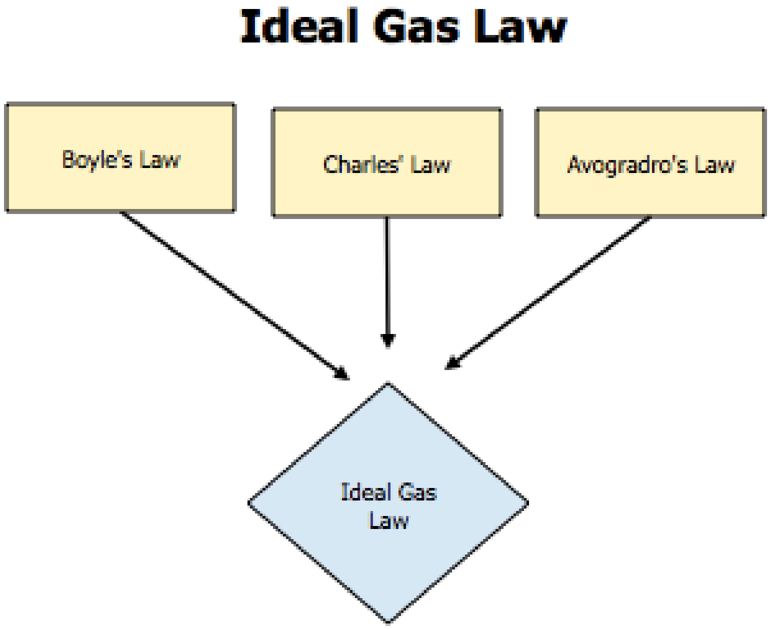
$$V = \frac{n_{N_2}RT}{P}$$

$$= \frac{2.09 \times 10^{-2}\ \cancel{mol} \times 0.08206\ \cancel{atm}\ L\cancel{mol^{-1}}\cancel{K^{-1}} \times 298.15\ \cancel{K}}{0.956\ \cancel{atm}}$$

$$= 0.535\ L$$

$$= 535\ mL\ N_2$$

Does an ideal gas exist ?



IDEAL AND NON-IDEAL BEHAVIOURS OF GASES

Ideal gas (Ideal behaviour)	Real gas (Non-ideal behaviour)
❖ Volume of gas particles is negligible compared to the volume of container.	❖ Gas particles have a certain volume & do occupy some space.
❖ Intermolecular forces between gas particles are negligible.	❖ There is intermolecular forces between gas particles

Deviation of gas from Ideal Behaviour

- Real gases do not show ideal behaviour.

- The deviation is due to:

  i) Gas particles occupy space, that is, they have volume.

  ii) The existence of intermolecular forces between gas particles .

The deviation is more significant at high pressures and low temperatures.

1) At high pressures:

- volume of container decreases
- gas particles are closer to each other and begin to occupy a space portion in the container.

Hence,

1. The volume of gas particles is significant.
2. The intermolecular forces between molecules become significant/important.

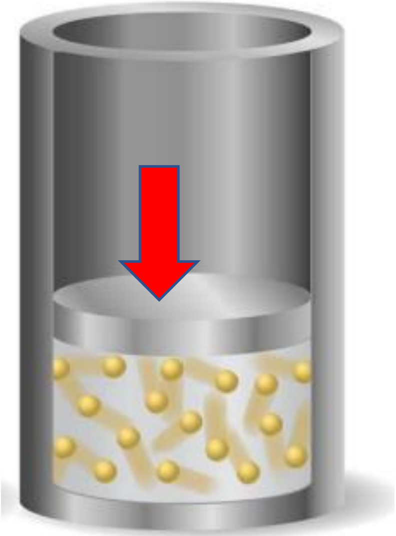
2) At low temperatures

- The kinetic energy of the gas particles decrease. Hence,

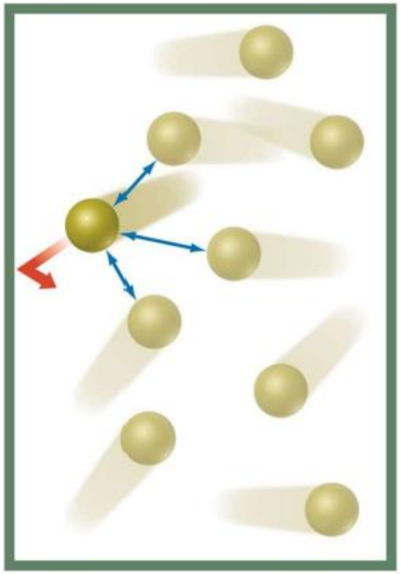
  i. The gas particles move at low speed

  ii. The intermolecular forces between molecules become significant.

Deviation from Ideal Behaviour at High Pressures



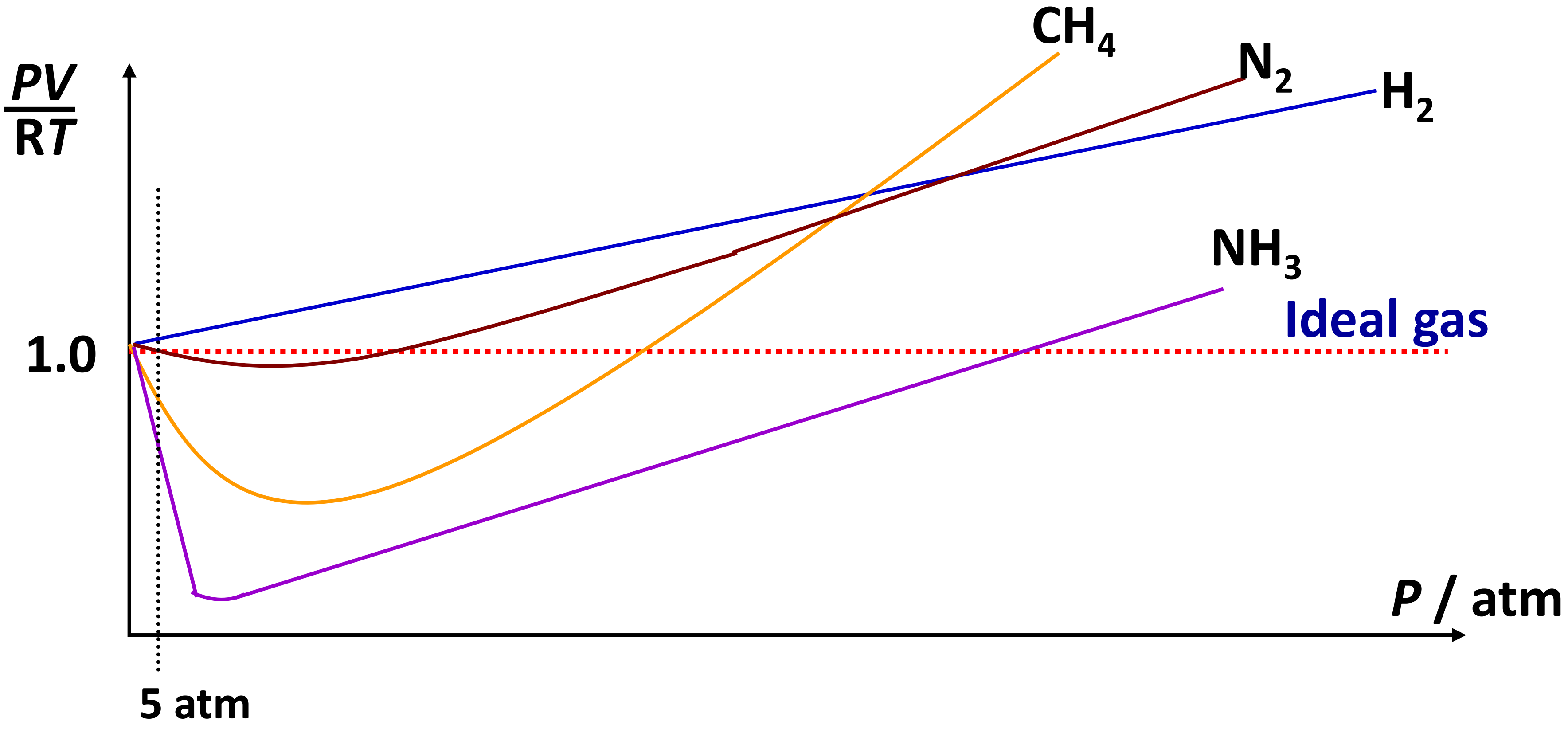
High pressure



Real gas

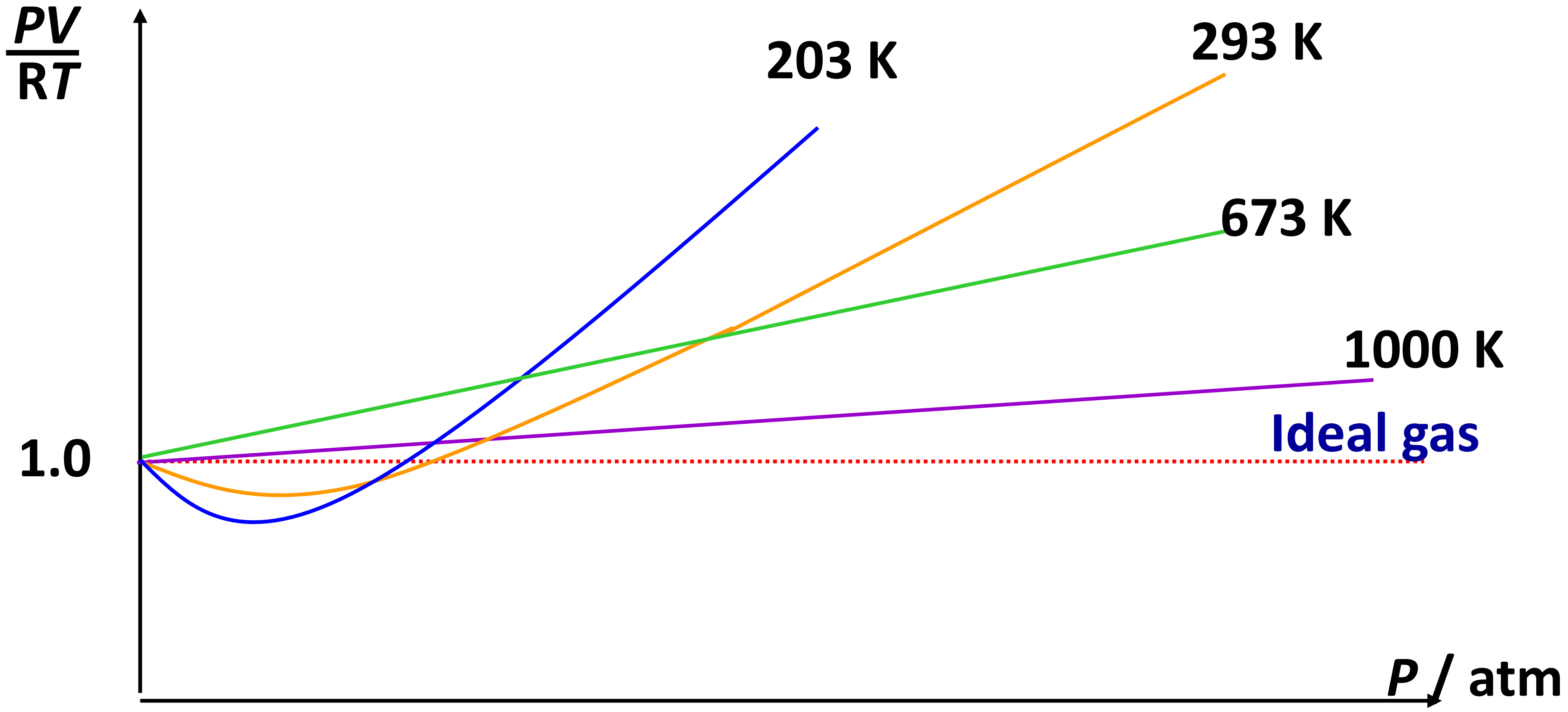
At high pressures, the particles are much closer together and the attractive forces cannot be ignored. These attractive forces will decrease the force of collisions between the gas molecules and walls of container, therefore reducing the pressure exerted compared to an ideal gas.

Plot of PV/RT versus the Pressure(P) of 1 mole of Various Gases at 273 K(0°C)



- When $P\uparrow$
- Deviation from ideal behavior ↑

Plot of PV/RT versus the Pressure(P) of 1 mole of N_2 at Various Temperature



- When T ↓ and P ↑
- Deviation from ideal behavior ↑

van der Waals Equation

- **To describe real gas behavior accurately, we need to modify the ideal gas equation → van der Waals equation**

- **van der Waals adjusted the ideal gas equation by correcting two parameters:**

  i) The volume of gas particles occupied in the container are considered.

  ii) The intermolecular forces acting between gas particles are considered.

van der Waals equation :

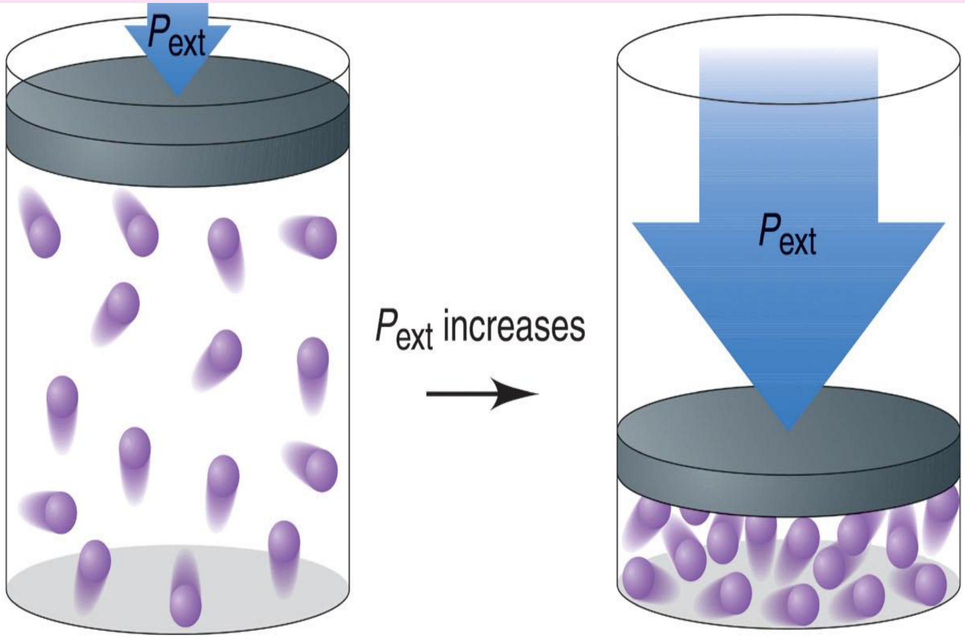
$$\left[P + \frac{n^2a}{V^2}\right](V - nb) = nRT$$

Correction for molecular attraction

Correction for molecular volume

Where, P = Pressure of gas
V = Volume of container
n = Mole of gas
a, b = van der Waals constant

(i) The effect of molecular volume on measured gas volume

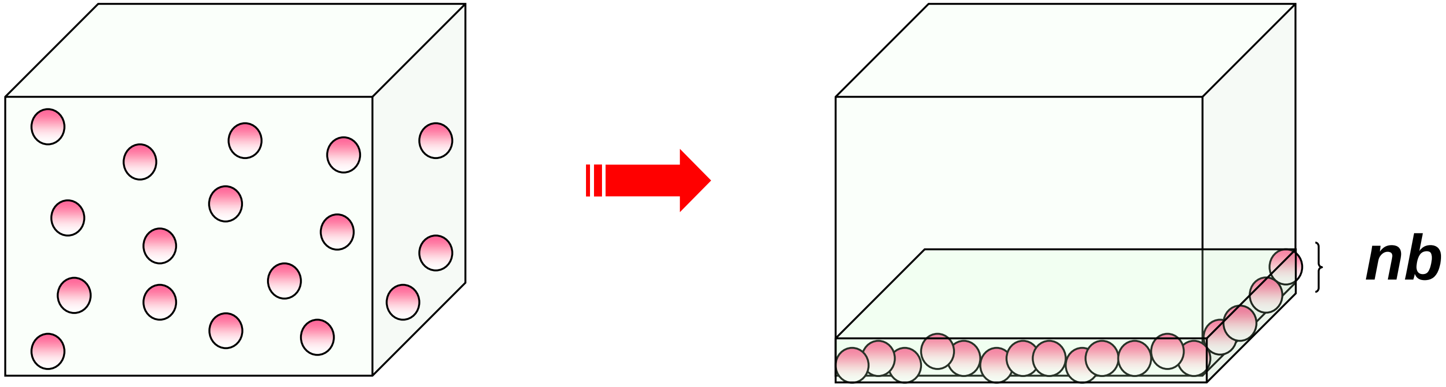


Deviation from Ideal behavior

At lower pressure, the volume of the gas particles is negligible compared to the total volume ($V_{gas} = V_{container}$) and the gas behave as an ideal gas.

At high pressure, the volume of the gas particles is more significant compared to the total volume. As a result, the volume of a real gas at high pressure is larger than the ideal gas.
($V_{real} > V_{ideal}$).

- **The increase in V of real gas is caused by the effect of molecular volume**
- **It is corrected by subtracting an amount *nb*, which accounts for the finite volume occupied by the gas particles.**



$$
\therefore \quad \begin{aligned} V_{\text{finite}} &= V_{\text{container}} - V_{gas} \\ &= V_{\text{container}} - nb \end{aligned}
$$

n **: moles of molecules**

b **: constant representing the volume occupied by the gas particles**

$$\begin{aligned} V_{\text{finite}} &= V_{\text{container}} - V_{gas} \\ &= V_{\text{container}} - nb \end{aligned}$$

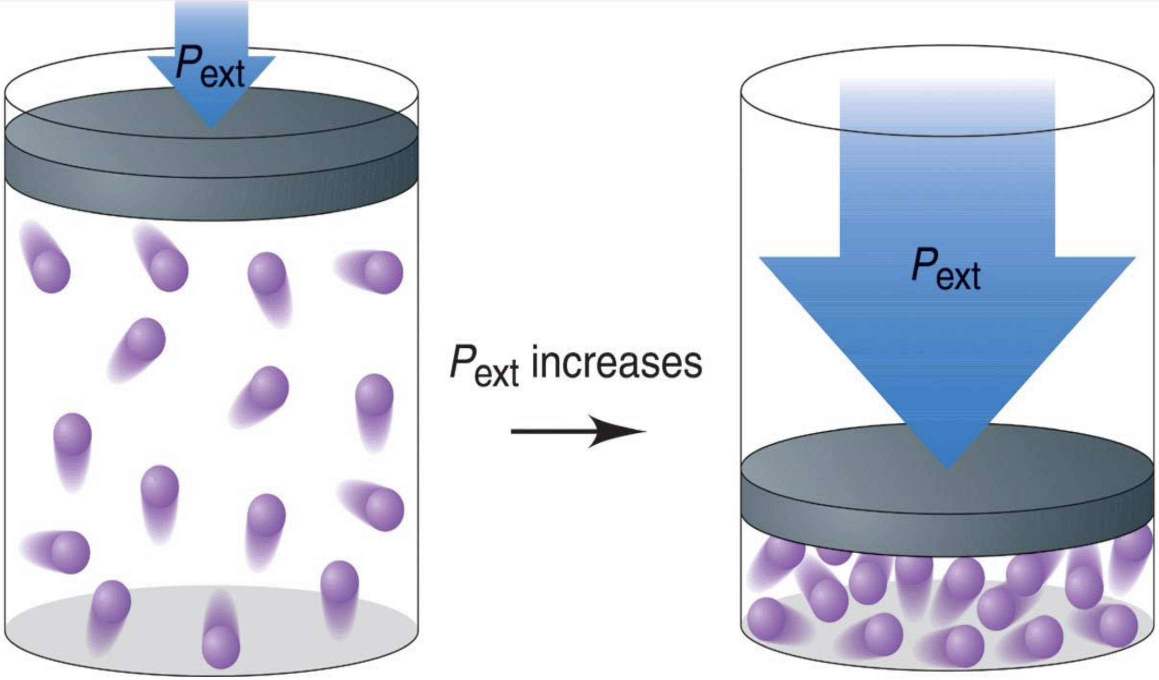
n : moles of molecules

b : constant representing the volume occupied by the gas particles

- **The value of b relates to molecular volume of the gas.**
- **The larger the molecule (or atom), the greater b is**

Molecule size ↑ , b↑

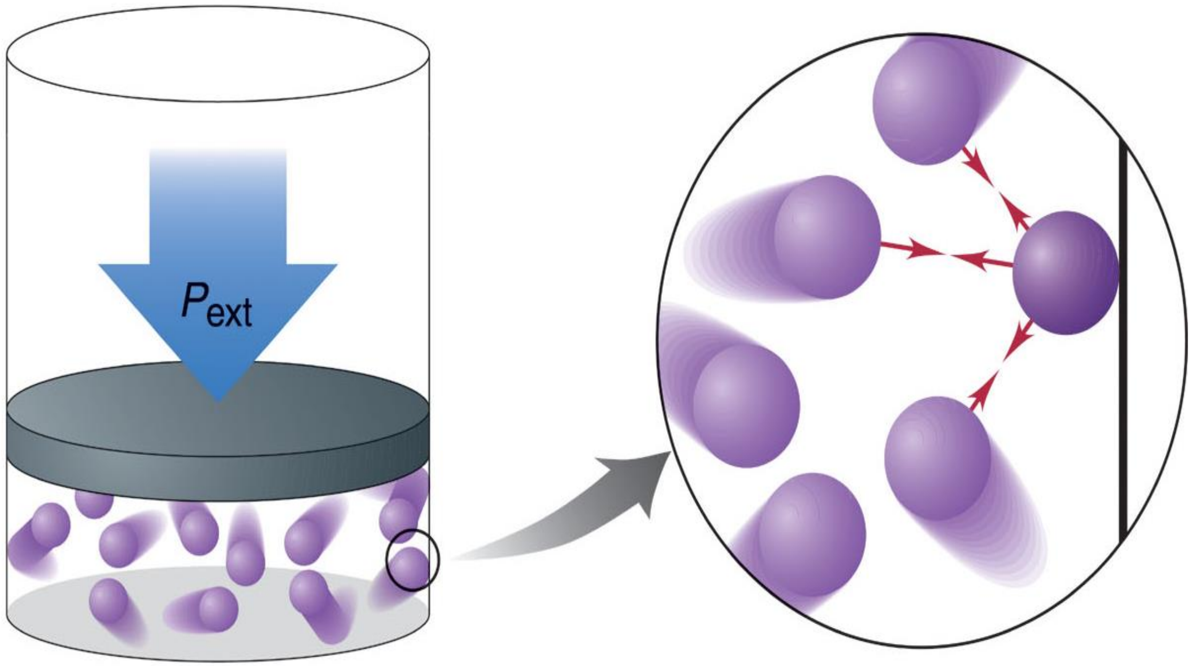
(ii) The effect of intermolecular forces on measured gas pressure



Deviation from Ideal behavior

At lower pressure, the gas particles are so fart apart. There are no attractive and repulsive forces between the gas particles. Therefore, the gas behave as an ideal gas.

At high pressure and low temperature, the gas particles are much closer together and move slower. The attractive and repulsive forces between them is significant. The particles give less impact to the wall. The pressure exerted by a real gas is lower than expected from ideal gas, $(P_{real} < P_{ideal})$



Moderately high P_{ext}:
molecules close
enough to interact

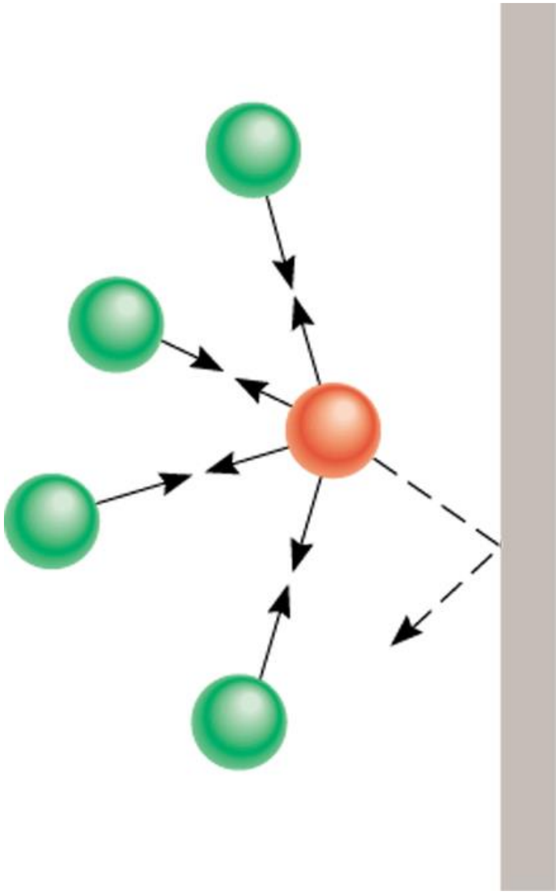
Attractions lower
force of collision
with wall

- The intermolecular forces have an effect on the speed of the moving gas particles.

- Gas particles that experience this force will move slowly.

- The collision of gas particles with the wall of container become less.

- lower pressure exert compare to ideal gas situation



$$\therefore \; P_{real} < P_{ideal}$$

- The decrease in P of real gas is caused by the effect of intermolecular attraction.

- It is corrected by adding an amount , $\frac{n^2a}{V^2}$ which accounts for the intermolecular forces between gas particles.

$$P_{\text{real}} < P_{\text{ideal}}$$

$$\therefore P_{\text{ideal}} = P_{\text{real}} + \frac{n^2a}{V^2}$$

n : moles of particles
a : constant relates to intermolecular forces between gas particles

The values of *a* indicates the strength of intermolecular forces acting between molecules.

intermolecular forces ↑ , *a* ↑

$$P_{\text{ideal}} > P_{\text{real}}$$

$$P_{\text{ideal}} = P_{\text{real}} + \frac{n^2a}{V^2}$$

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

$$V = V_{\text{container}} - V_{gas}$$
$$= V_{\text{container}} - nb$$

Ideal Gas equation : $PV = nRT$

van der Waals equation :

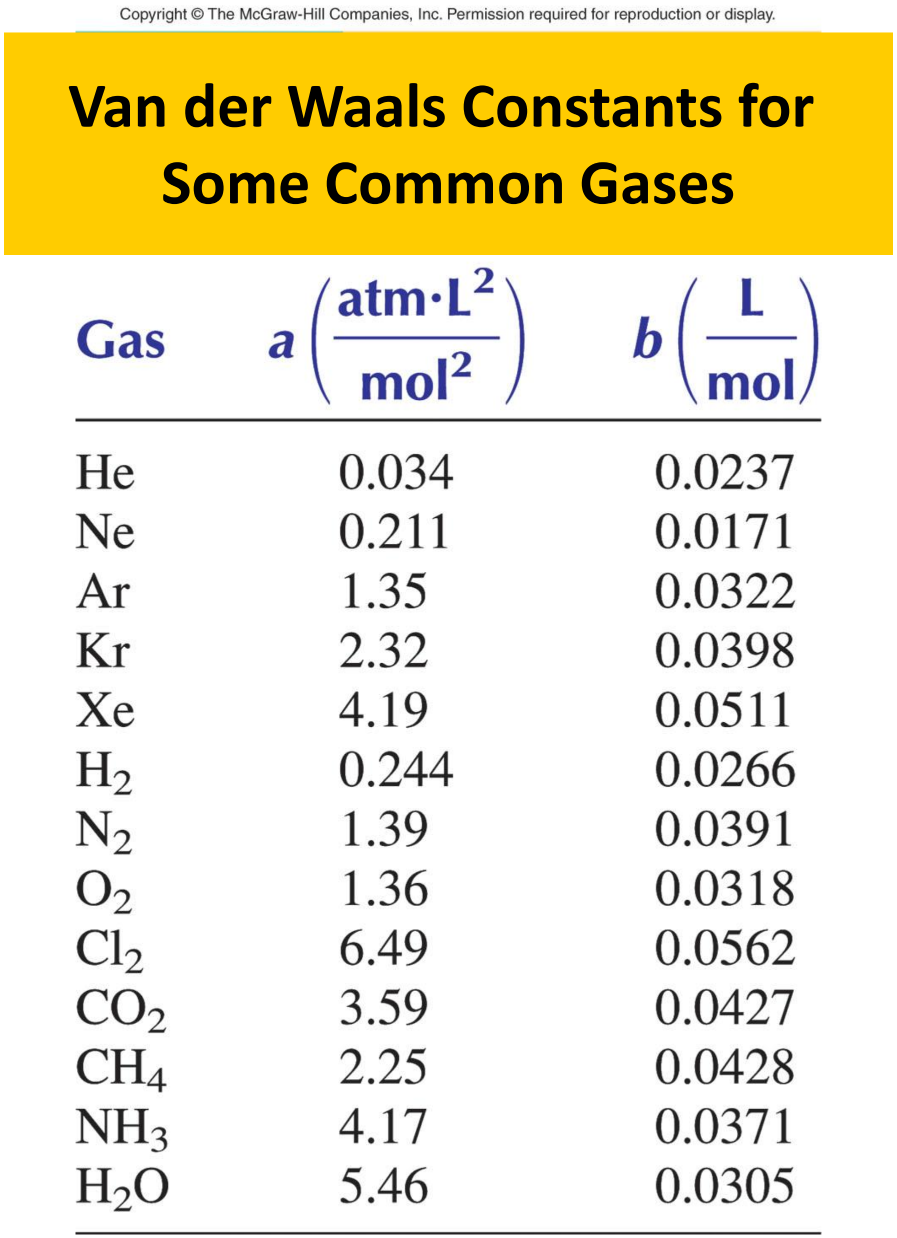
$$\left[P + \frac{n^2a}{V^2}\right](V - nb) = nRT$$

Correction for molecular attraction

Correction for molecular volume

Where, P = Pressure of gas
V = Volume of container
n = Mole of gas
a,b = van der Waals constant

The constant value a and b are different for different gases



Van der Waals Constants for Some Common Gases

Gas	$a\left(\frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2}\right)$	$b\left(\frac{\text{L}}{\text{mol}}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.35	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0511
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
NH_3	4.17	0.0371
H_2O	5.46	0.0305

CONDITIONS AT WHICH REAL GASES APPROACH THE IDEAL BEHAVIOUR

❶ At very low pressure

(Constant temperature)

- Volume of container is increased.
- The gas particles are far apart in a large volume
- The intermolecular forces are negligible
- Volume of gas molecules is too small compared to volume of the container, thus volume of gas molecules are negligible.

Hence, it obey kinetic–molecular theory, the gas behave ideally.

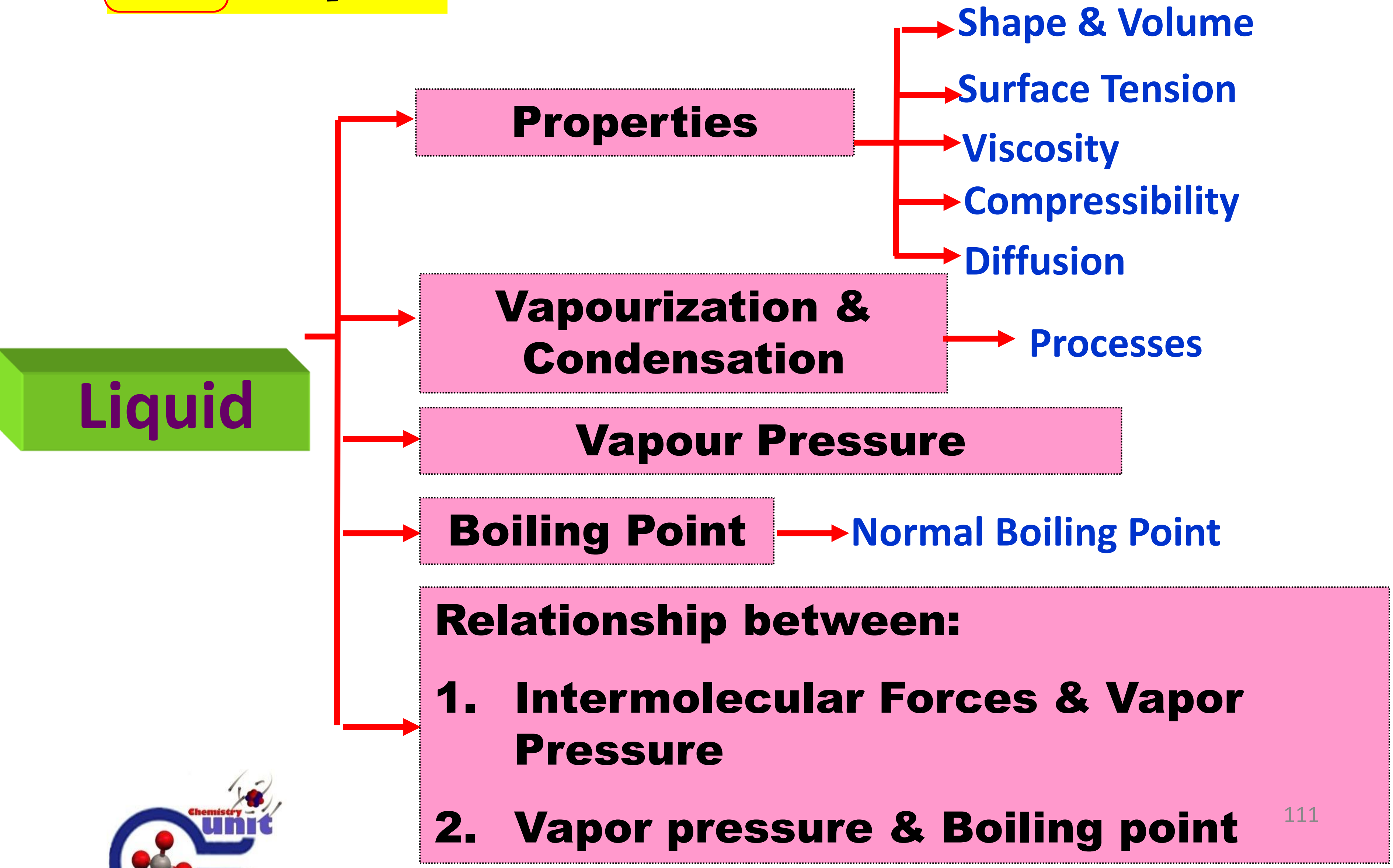
❷ At high temperature

(Constant volume)

- Gas particles have high kinetic energy and move at high speed.
- Gas particles are free from intermolecular forces that act between them.
  - intermolecular forces are negligible

Hence, it obey kinetic–molecular theory, the gas behave ideally.

5.2 Liquid



Liquid

- **Properties**
  - Shape & Volume
  - Surface Tension
  - Viscosity
  - Compressibility
  - Diffusion
- **Vapourization & Condensation**
  - Processes
- **Vapour Pressure**
- **Boiling Point**
  - Normal Boiling Point
- **Relationship between:**
  1. Intermolecular Forces & Vapor Pressure
  2. Vapor pressure & Boiling point

5.2 LIQUIDS

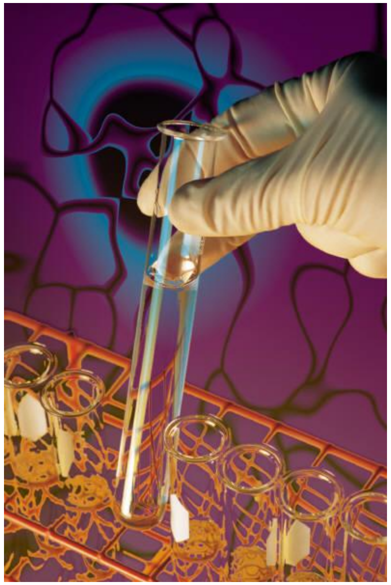
Teaching and learning outcomes

At the end of the lesson, student should be able to

5.2	Liquids
a)	Relate the properties of liquid to intermolecular forces, molecular arrangement and molecular forces, molecular arrangement and molecular motion in explaining shape, volume, surface tension, viscosity, compressibility and diffusion. (C2, C3)
b)	Explain vaporisation and condensation processes based on kinetic molecular theory and intermolecular forces. (C2, C3)
c)	Define: (C1) i. vapour pressure ii. boling point.
d)	Explain boiling process. (C2, C3)
e)	Illustrate the relationship between: (C2, C3, C4) i. Intermolecular forces and vapour pressure ii. Vapour pressure and temperature

PROPERTIES OF LIQUID

- Conforms to shape of container
- Volume limited by surface
- Only slightly compressible
- Moderate ability to flow
- Molecules slide past one another freely



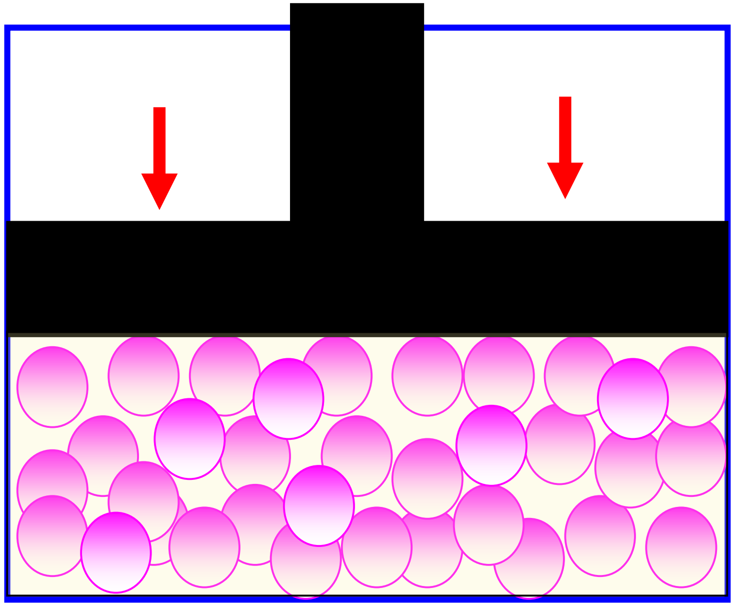
SHAPE AND FLUIDITY

- Liquid has a definite volume but not a definite shape
- Molecules arranged closely but not rigidly
- Molecules held together by intermolecular force, but they able to move freely
- **Moderately highly fluidity and take the shape of container**
- **Liquid conforms to shape and volume of container**

COMPRESSIBILITY

- Very little free space between molecules

- Compress only slightly & more difficult to compress than gas



DIFFUSION

- Liquid diffuse slower than gas

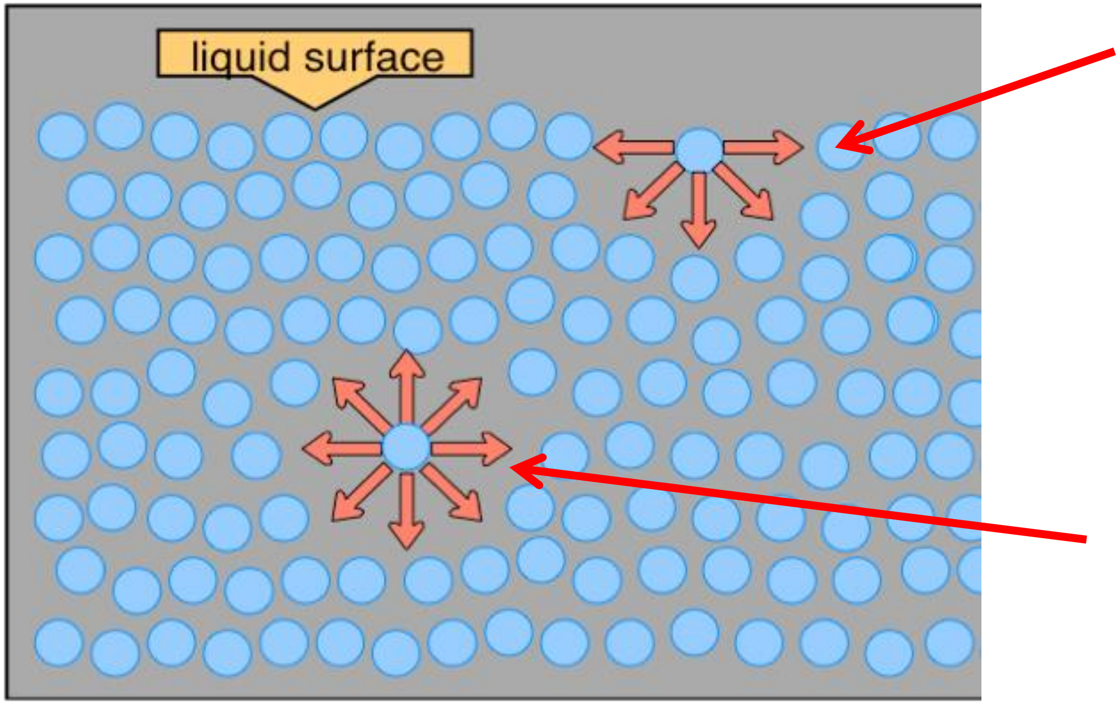
Stronger intermolecular force and closer molecular arrangement

SURFACE TENSION

- Surface tension is the energy required to increase surface area by a unit area

A paper clip will float on water if it is carefully placed on the surface of the water. It is held up by surface tension

- It is caused by the difference in intermolecular forces experienced by molecules at the surface of a liquid and those experienced by molecules in the interior



Molecules at the surface feel attractive forces from below and sideways but not upwards. Thus they are pulled in toward the liquid.

Molecules in the interior are surrounded and are pulled equally in all directions.

- **Therefore, a liquid surface tends to have the smallest possible area.**

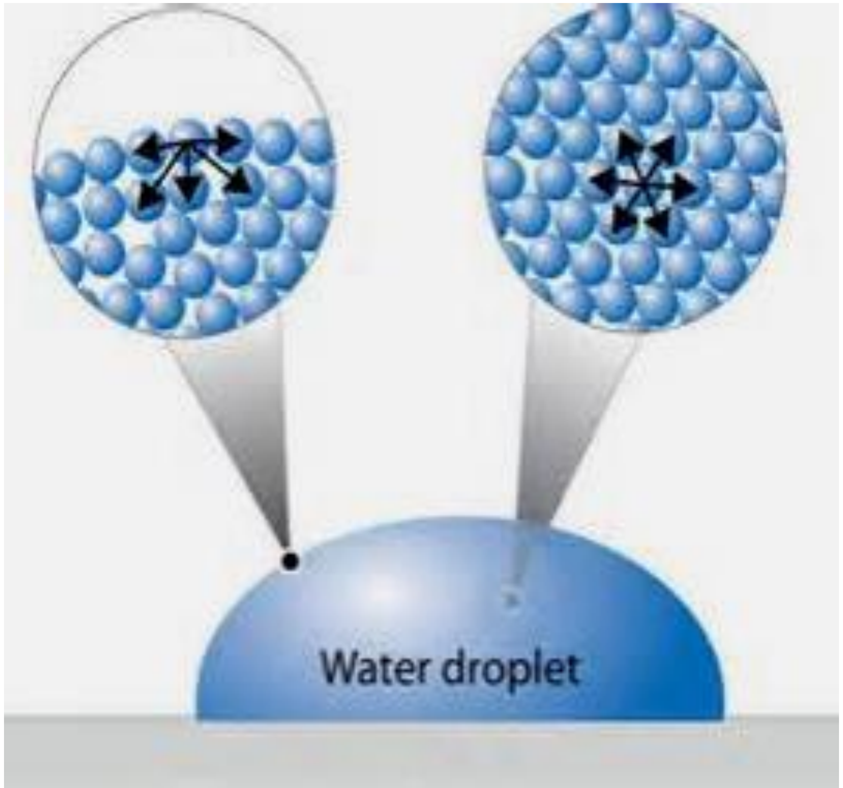
- **To increase the surface area, molecules must move to the surface by breaking some attractions in interior, which requires energy and cause the surface tighten like an elastic film**

FACTOR AFFECTING SURFACE AREA:

Intermolecular forces ↑

surface tension ↑

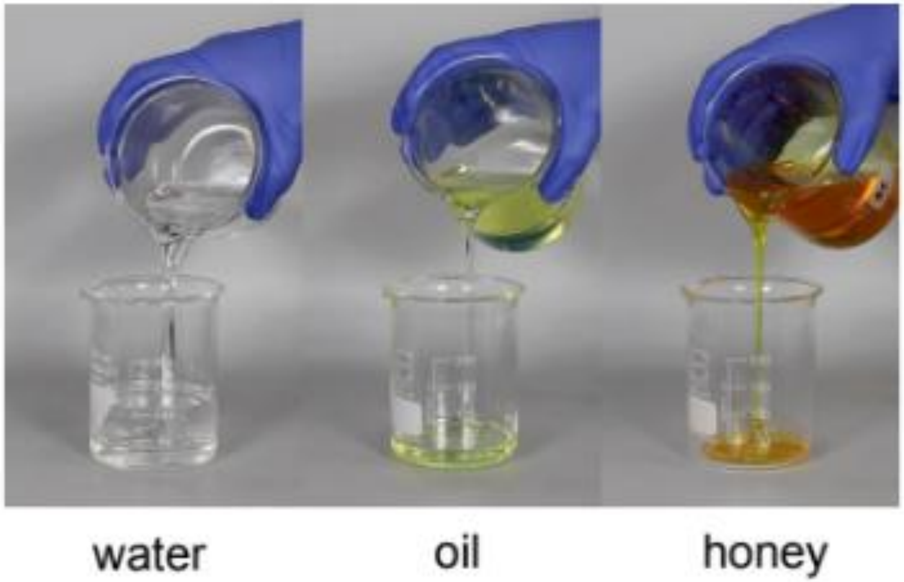
The stronger the forces are between molecules in a liquid, the greater the surface tension



VISCOSITY

Viscosity = resistance of the liquid to flow due to it's intermolecular forces.

- **Example of high viscosity liquid : ketchup, yogurt, honey**
- **Example of low viscosity liquid : oil, water, juice**



FACTOR AFFECTING VISCOSITY

i. **Intermolecular forces**

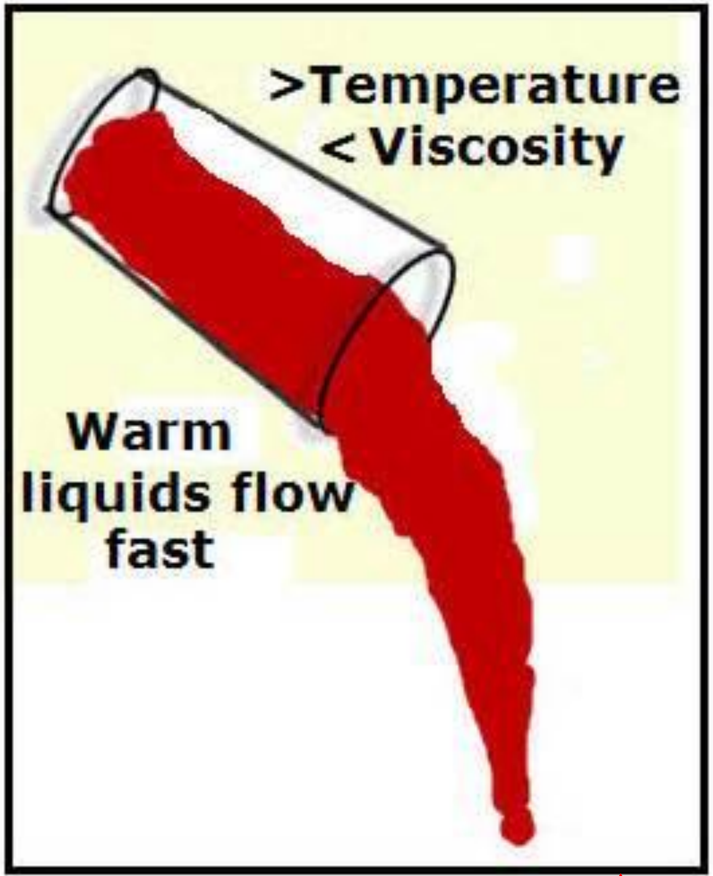
- Intermolecular forces stronger↑, resistance to flow ↑, viscosity ↑

ii. **Temperature**

- Temperature ↑, average kinetic energy of molecules ↑ , resistance to flow ↓, viscosity ↓

iii. **Size and shape of molecules**

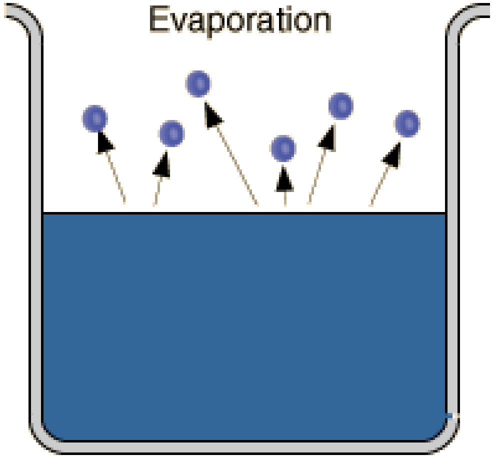
- Size of liquid molecules↑, resistance to flow ↑ ,viscosity ↑.



Temperature ↑
Viscosity ↓

Small and spherical molecules are easy to flow
(weaker intermolecular forces)
→ viscosity decrease

VAPORISATION AND CONDENSATION



VAPORIZATION	CONDENSATION
▪ Process in which a liquid is changed into a gas	▪ **Process in which a gas is changed into a liquid**
▪ When the kinetic energy of liquid molecules increases, - molecules will move faster. - more molecules possess sufficient kinetic energy to overcome the intermolecular forces **Molecules escape from the liquid surface as vapour molecules.**	▪ When the vapor molecules are cooled and lose kinetic energy, - molecules will move slower. - Kinetic energy of molecules can no longer overcome the intermolecular forces. - Molecules come closer, attract one another and return as liquid molecules.

FACTORS AFFECTING RATE OF VAPORIZATION

Surface area

Surface area ↑, vaporization rate ↑

- number of molecules with high energy have chance to escape from surface ↑

Temperature

Temperature ↑, vaporization rate ↑

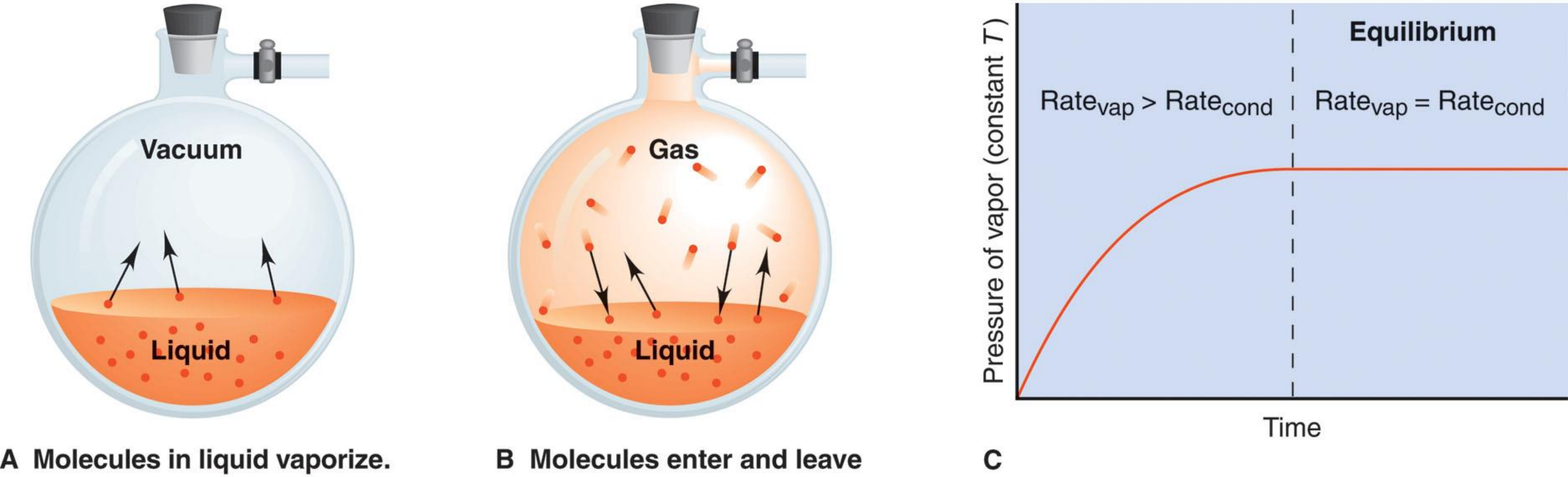
- more molecules have higher kinetic energy
- number of molecules escape from surface ↑

Strength of intermolecular forces

Strength ↓ , vaporization rate ↑

VAPOUR PRESSURE

- Definition: The pressure exerted by vapour in equilibrium with its liquid in a closed container.



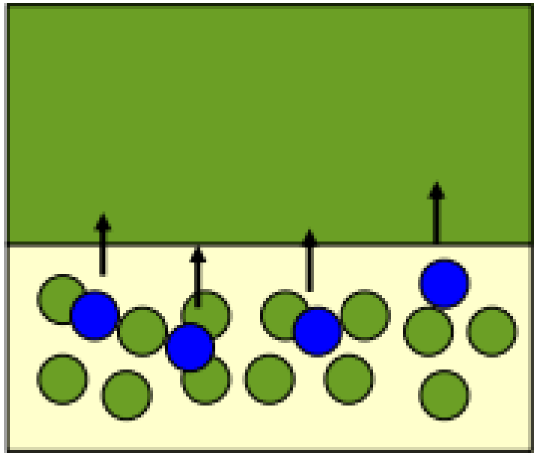
A Molecules in liquid vaporize.

B Molecules enter and leave liquid at same rate.

C

VAPOUR MOLECULES IN A CLOSED SYSTEM:

- They remain in the container and moves in constant random motions.



- They continually strike :
  - ✓ the wall of the container
  - ✓ other vapour molecules
  - ✓ the surface of the liquid molecules

- Every time the vapour molecules hit the wall of the container, pressure is exerted.

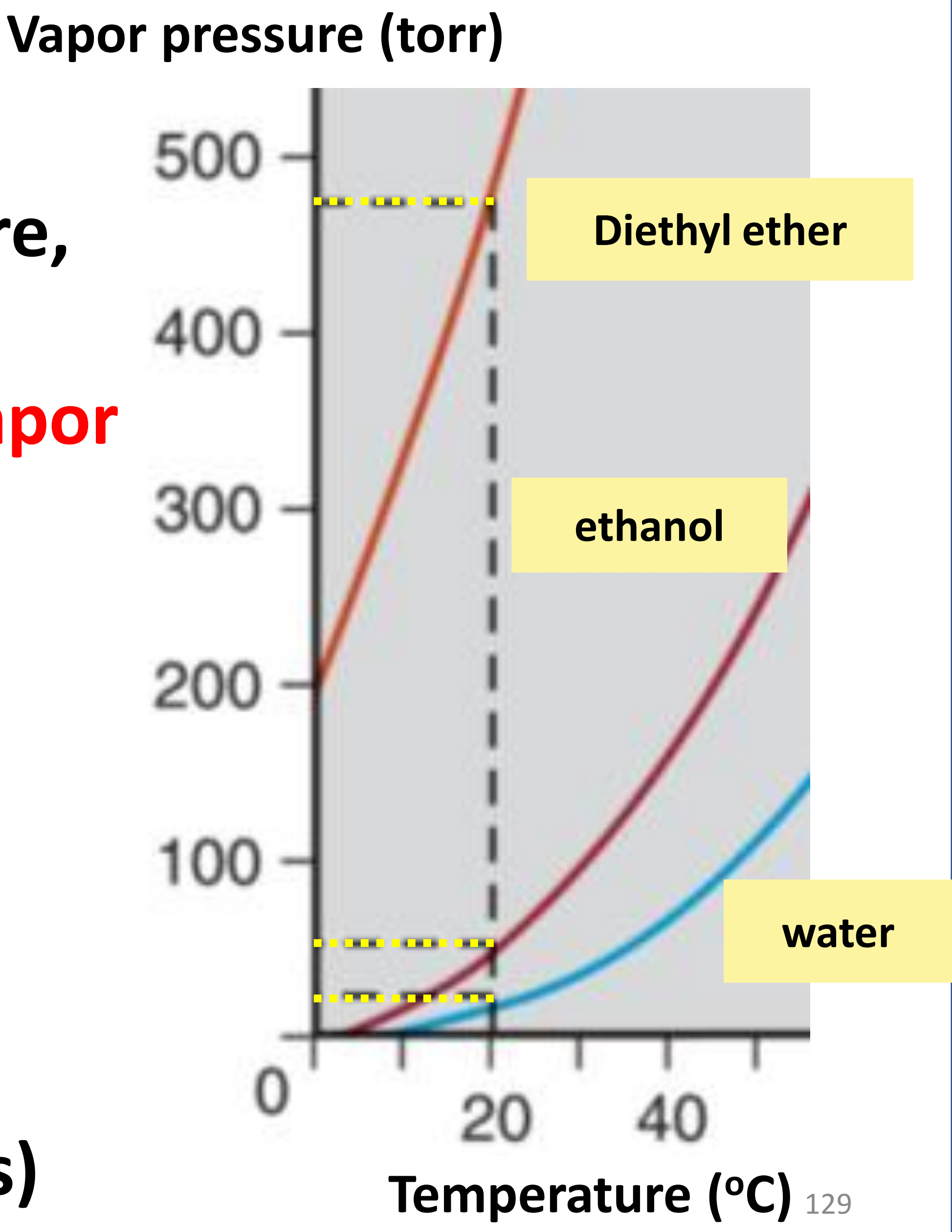
VOLATILITY OF LIQUID

- Volatile liquids
  - High vapor pressures at room temperature
- Nonvolatile liquids
  - Very low vapor pressures at room temperature

EXAMPLE:

- **At any given temperature, (example: 20^{o}C) diethyl ether has the highest vapor pressure and water the lowest**

- **iethyl ether has the weakest intermolecular (Van der Waals fores) forces and water the strongest (hydrogen bonds)**



FACTOR AFFECTING VAPOUR PRESSURE

Temperature

Temperature ↑

- Liquid molecules move faster,

Kinetic energy of molecules ↑

- easy to escape to form vapour molecules.

∴ vapour pressure is high

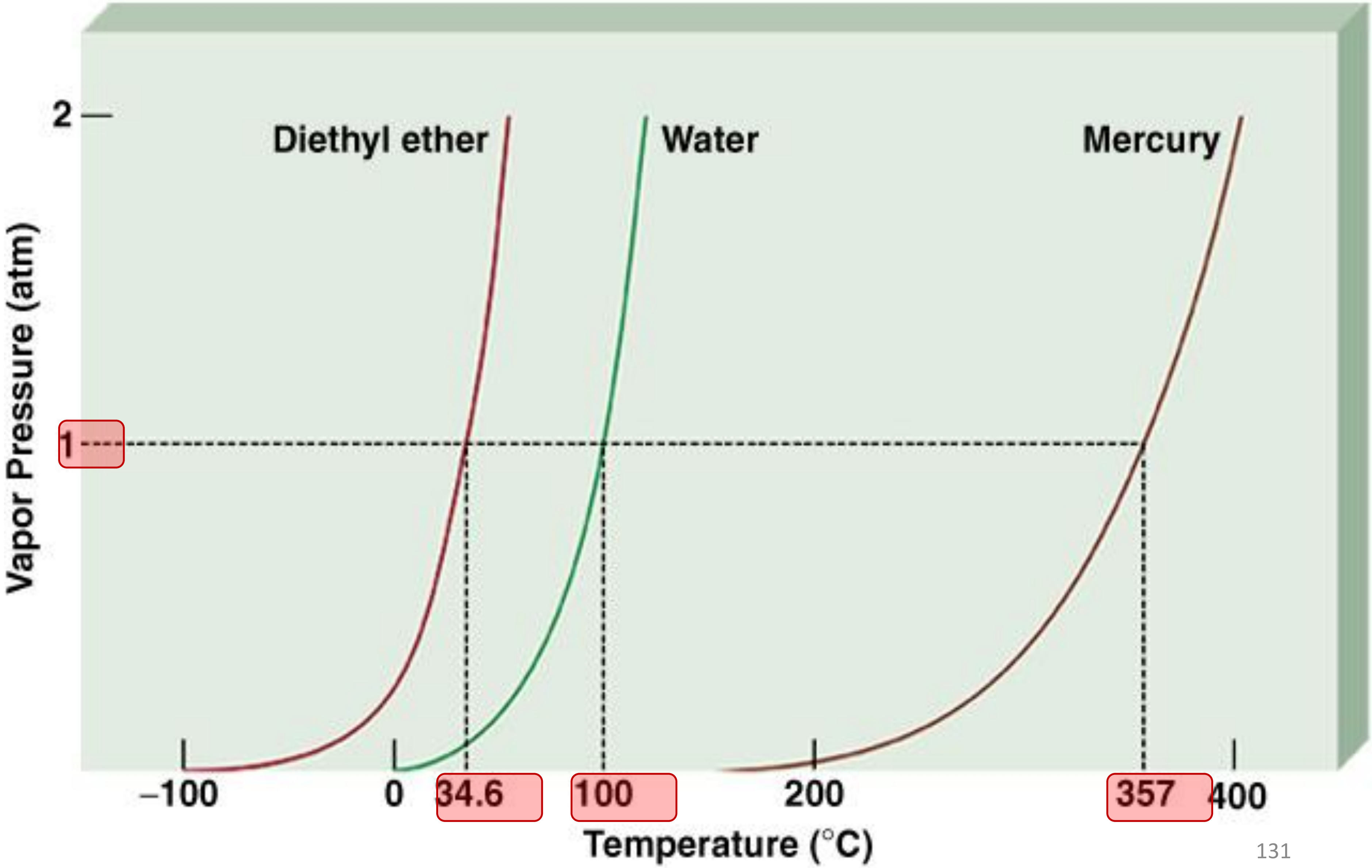
Strength of intermolecular forces

Weak intermolecular forces,

- liquid molecules easy to escape to form vapour.

∴ vapour pressure is high

Vapor Pressure vs. Temperature



BOILING POINT

- The temperature at which the vapour pressure of a liquid is equal to the external (atmospheric) pressure

EXAMPLE: boiling point of water

100^{o}C at 760 torr (sea level)

94^{o}C at 610 torr (Boulder, Colorado)

72^{o}C at 270 torr (Mt. Averest)

NORMAL BOILING POINT

- **The temperature at which a liquid boils** when the external pressure is 1 atm (standard atmospheric pressure)

KEEP IN MIND

- If the pressure is not mentioned when the boiling point is measured, we assume it as normal boiling point (at 1 atm) .

BOILING PROCESS

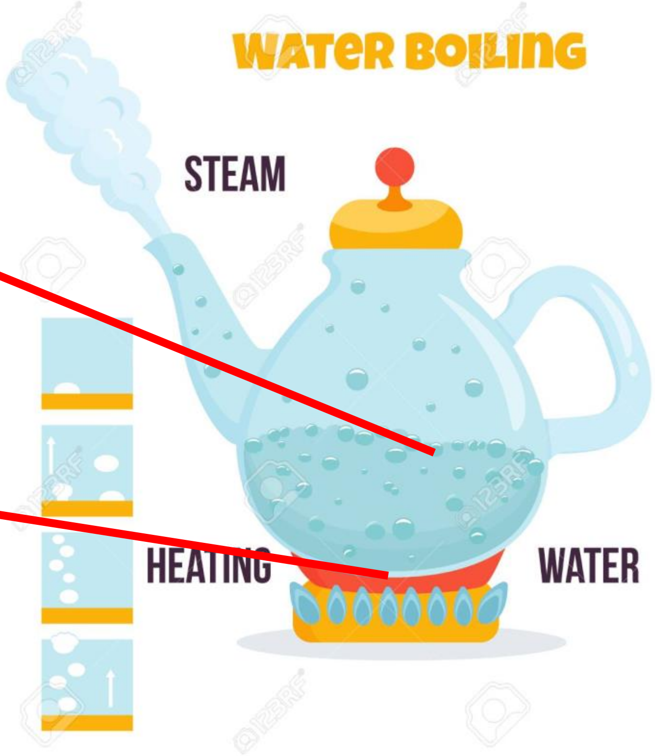
- As a liquid is heated, its vapor pressure increases.

- Continuing to increase the temperature will increase its vapor pressure until the vapor pressure of the liquid is equal to the external (atmospheric) pressure. At this stage, boiling starts.

- At this point, the changes of the state from liquid to gaseous does not only occur at the surface of the liquid but also the inner part of the liquid.

- In order to form vapor, the molecules of the liquid must overcome the intermolecular forces between molecules.

- ❑ **We will observe bubbles forming within the liquid and therefore the boiling process occur. Bubbles only contain water molecules in the form of vapor (gas).**



- ☐ **The temperature at which the liquid boils (temperature at which the vapor pressure of a liquid is equal to atmospheric pressure) is called the boiling point.**

- ☐ **Since the vapor pressure remains constant, the temperature will remain the same during the process of boiling even more heat is supplied.**

FACTOR AFFECTING BOILING POINT

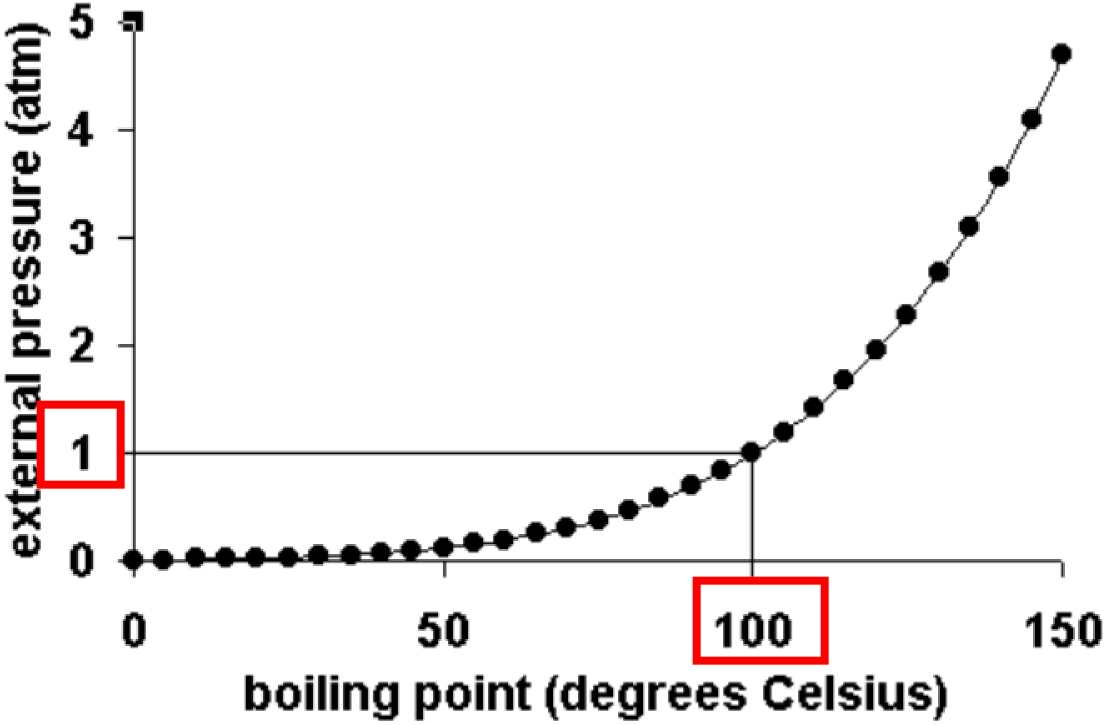
1) Atmospheric pressure

- When the atmospheric pressure ↓
- Heat needed to make vapor pressure of a liquid equals to atmospheric pressure ↓
- Boiling point ↓

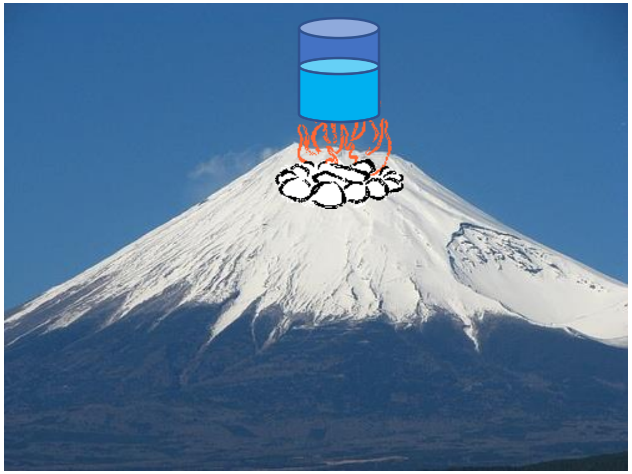
Boiling point $\propto$ Atmospheric pressure

The following graph shows the boiling point for water as a function of the external (atmospheric) pressure.

The line on the graph shows the normal boiling point for water.



Boiling point of water



Mt. Fuji
(3776 m)
Bp = 87°C

Mt. Everest
(8848 m)
Bp = 71°C

2) Intermolecular forces

- **Strength of intermolecular ↑ , boiling point ↑**

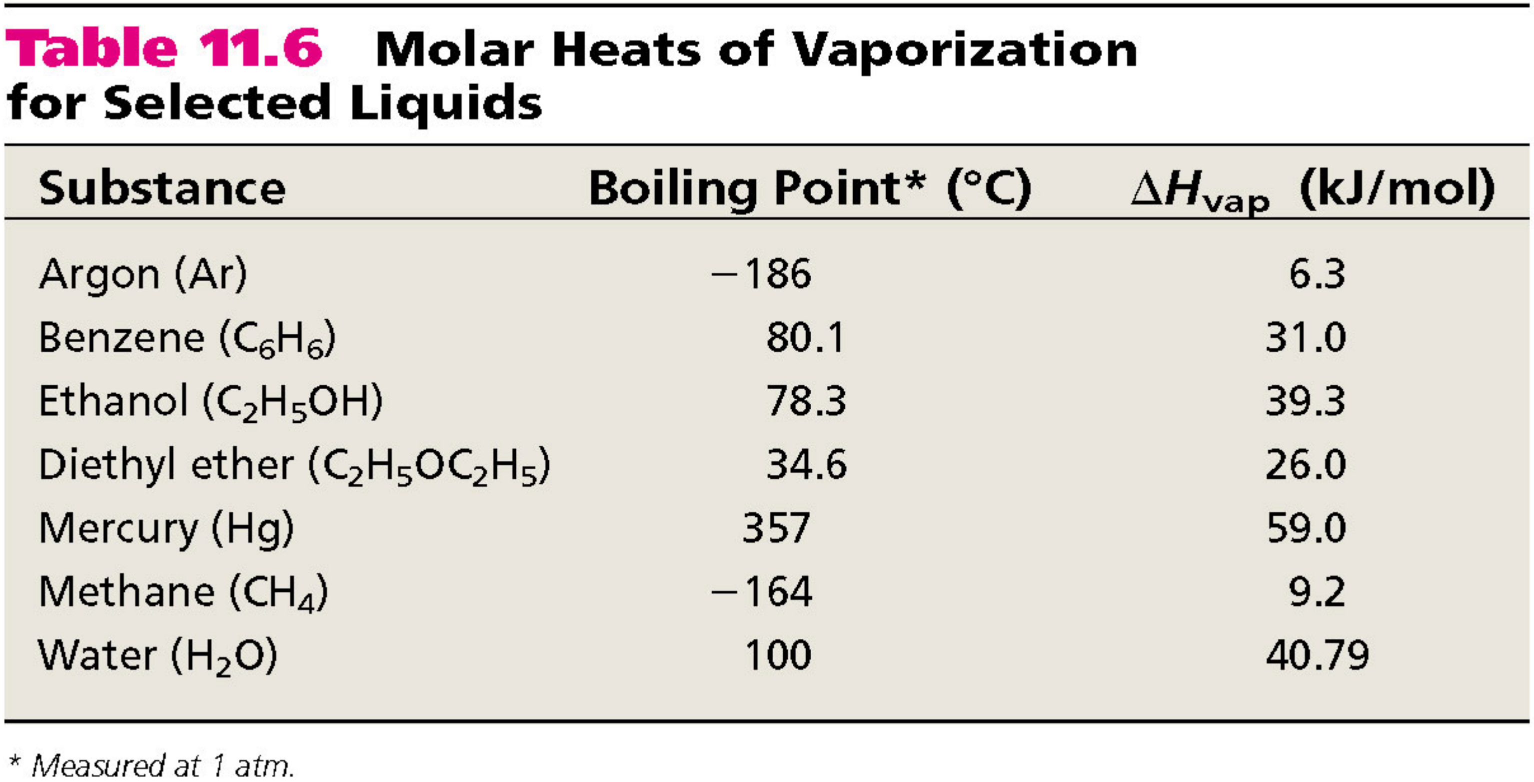


Table 11.6 **Molar Heats of Vaporization for Selected Liquids**

Substance	Boiling Point* (°C)	ΔH_{vap} (kJ/mol)
Argon (Ar)	−186	6.3
Benzene (C_6H_6)	80.1	31.0
Ethanol (C_2H_5OH)	78.3	39.3
Diethyl ether ($C_2H_5OC_2H_5$)	34.6	26.0
Mercury (Hg)	357	59.0
Methane (CH_4)	−164	9.2
Water (H_2O)	100	40.79

* *Measured at 1 atm.*

KEEP IN MIND

Strength of intermolecular forces depend on:

- Types of forces

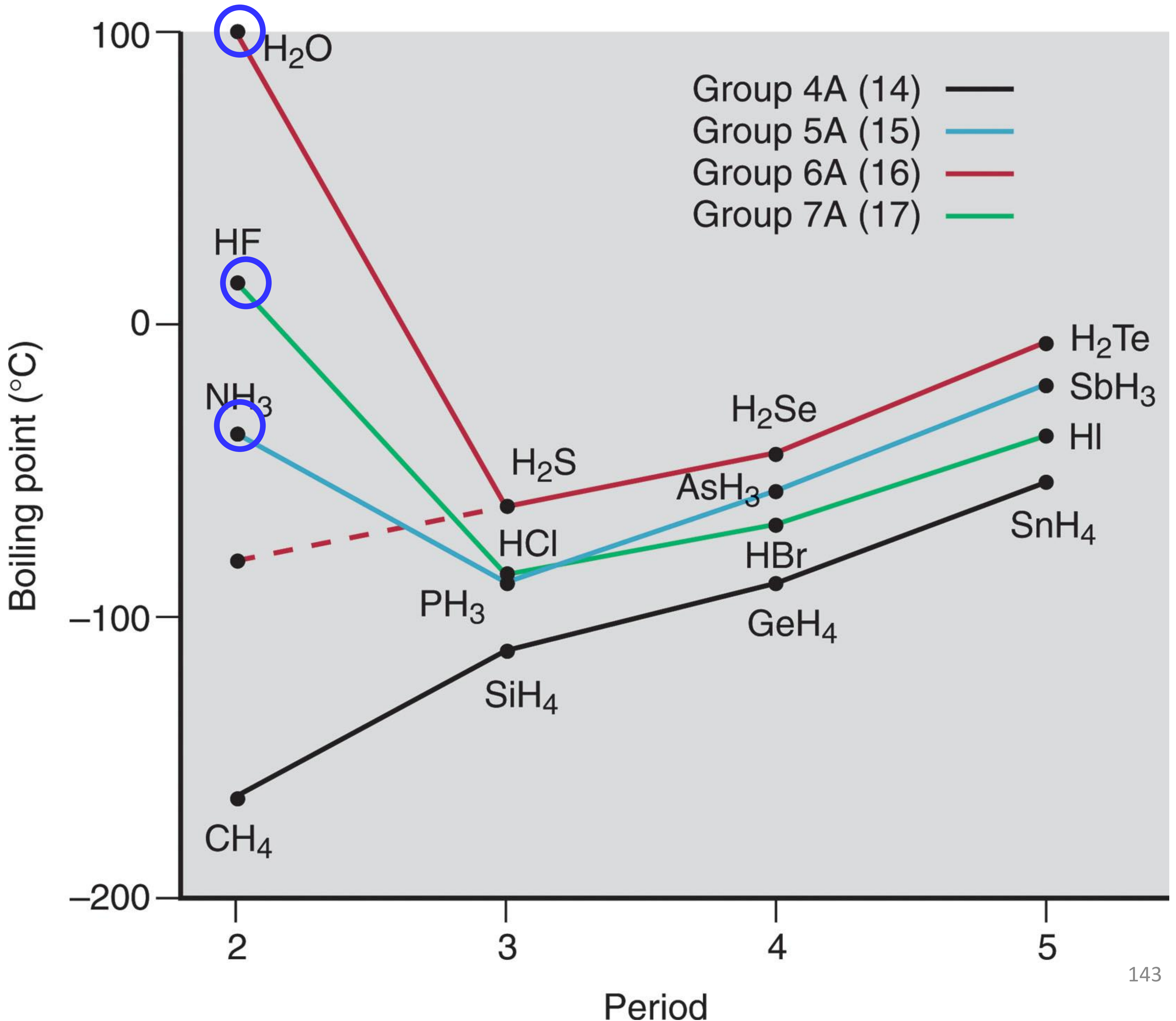
EXAMPLE:

Hydrogen bond is much stronger than Van der Waals force

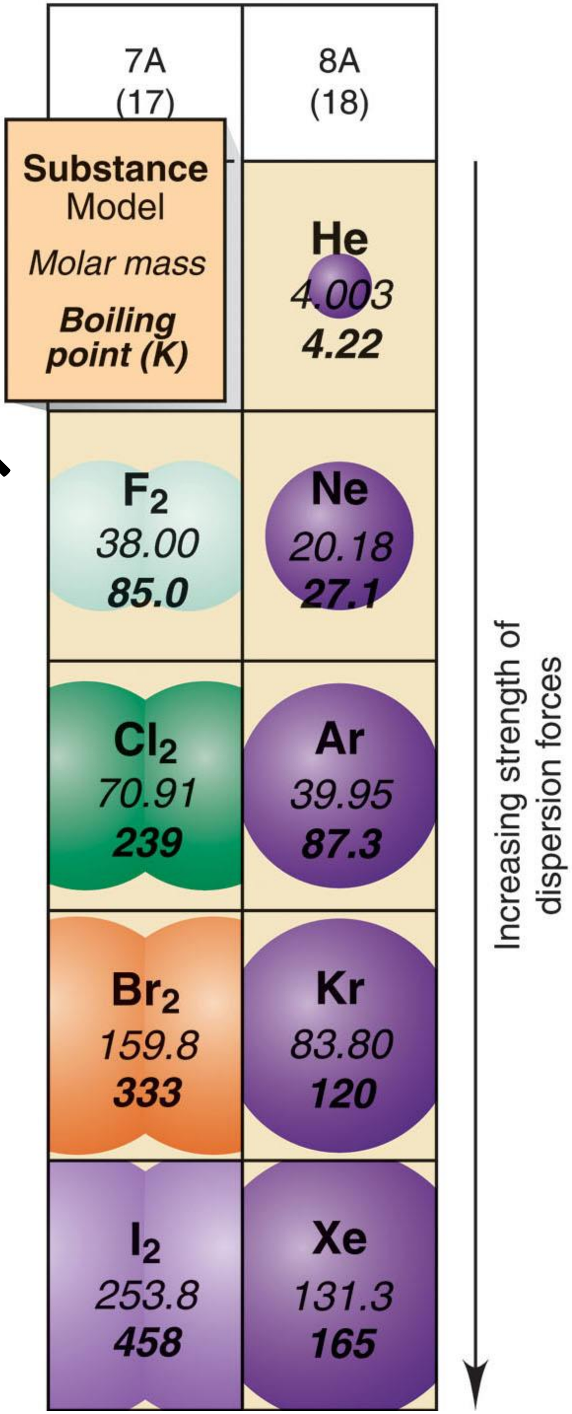
- Molar mass

EXAMPLE:

Cl_2 (molar mass = 70.91 g/mol) has a higher boiling point than F_2 (molar mass = 38.00 g mol^{-1})


100
0
–100
–200
Boiling point (°C)
2
3
4
5
Period
Group 4A (14)
Group 5A (15)
Group 6A (16)
Group 7A (17)
H_2O
HF
NH_3
CH_4
H_2S
HCl
PH_3
SiH_4
H_2Se
AsH_3
HBr
GeH_4
H_2Te
SbH_3
HI
SnH_4

- Size of atom / molecule ↑
- Intermolecular forces ↑
- Vapour pressure ↓
- Boiling point ↑



SUMMARY

weaker attractive force

higher vapor pressure

Lower boiling point

Liquids that have weaker attractive forces will have higher vapor pressure and lower boiling point.

**The more volatile a liquid, the lower will be its boiling point, the higher will be its vapor pressure.*

VAPORISATION VS BOILING PROCESS

VAPORIZATION	BOILING
▪ Process by which liquid changes into its vapors at temperature below its boiling point	▪ Process by which liquid changes to its vapors at its boiling point
• Takes place at all temperatures	• Takes place at only fixed temperatures
• Temperature may changes during vapourisation	• Temperature during boiling does not change
• A slow and steady process	• A fast and violent process
• Vapourisation takes place only at surface of liquids	• Boiling takes place throughout entire liquid

5.3 Solid

Solid

- **Explain the process of:**
  - Freezing (solidification)
  - Melting (fusion)
  - sublimation
  - deposition
- **Differentiate between amorphous and crystalline solids**
- **Types of crystalline solids with appropriate examples.**
  - Metallic
  - Ionic
  - Molecular covalent
  - Giant covalent

5.3 SOLIDS

Teaching and learning outcomes

At the end of the lesson, student should be able to

5.3	Solids
a)	Explain the following process using kinetic molecular theory: (C3) i. freezing (solidification); ii. melting (fusion); iii. sublimation; and iv. deposition.
b)	Differentiate between amorphous and crystalline solids. (C2)
c)	State the following types of crystalline solids with appropriate examples: (C1) i. Metallic; ii. Ionic; iii. Molecular covalent; iv. Giant covalent

PROPERTIES OF SOLID

Why solid are in fixed-shape ??

- Atoms / molecules in solids are arranged closely together & definite order.
- They can only vibrate and rotate about fixed position.
- They have rigid arrangement & cannot move freely

PROPERTIES OF SOLID

Shape and volume	Maintain it own shape and size
Compressibility	Almost none
Ability to flow	Almost none
Density	High
Motion of molecules	Vibrate about fixed positions

FREEZING (SOLIDIFICATION)

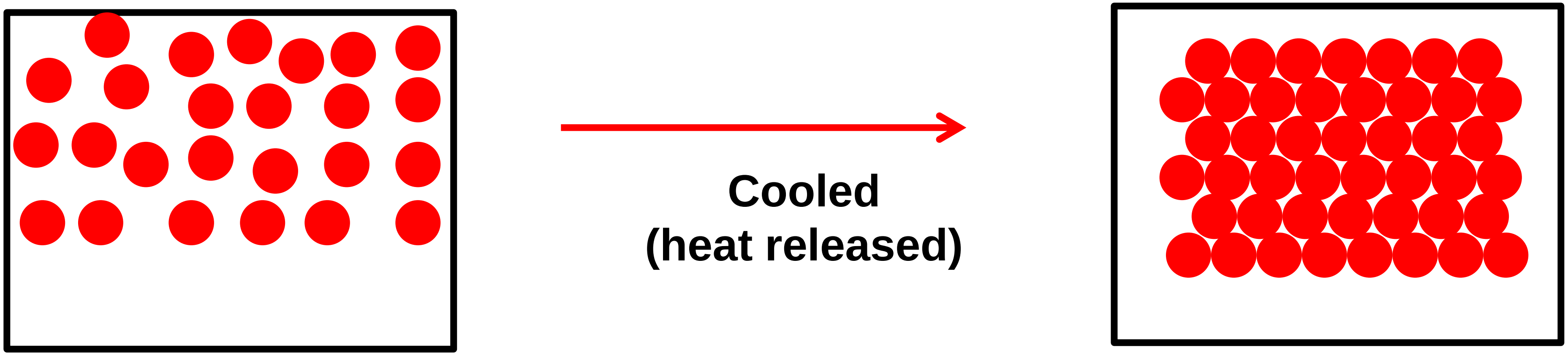
- A process of liquid changes to solid

EXAMPLE:

$H_2O(l) \rightarrow H_2O(s)$

- When liquid cooled, the particles lose kinetic energy and move more slowly.

∴ Intermolecular forces become stronger and particles become fixed in position.



Cooled
(heat released)

liquid **solid**

As the temperature drops, fewer particles have enough kinetic energy to overcome neighbouring attractions, and the particles lock into place. As more and more of the particles settle into a relatively fixed position, the substance gains a shape of its own and becomes a solid.

MELTING (FUSION)

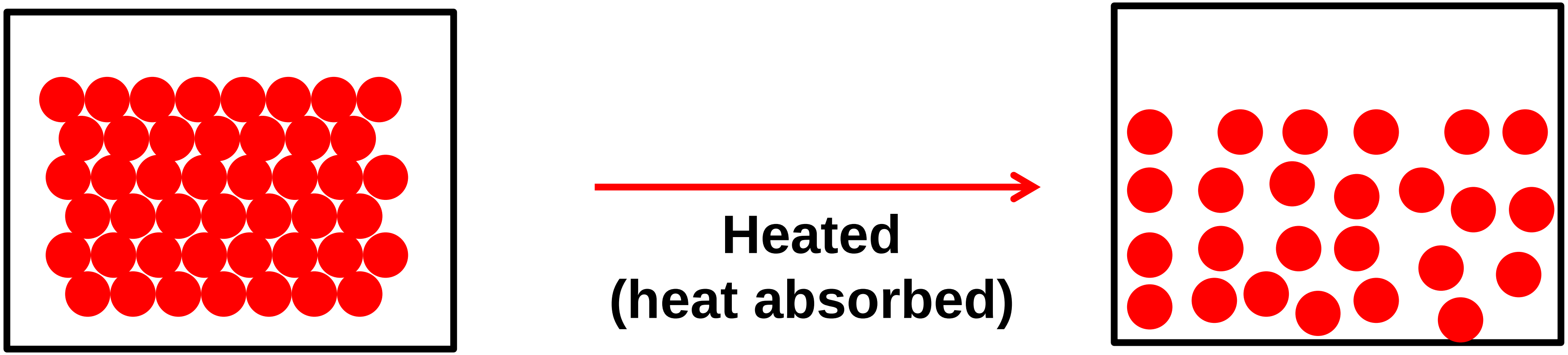
- A process of solid changes to liquid

EXAMPLE:

$H_2O(s) \rightarrow H_2O(l)$

- When solid heated, the particles gain kinetic energy and vibrate faster

- At certain temperature, the particles have enough kinetic energy to break free of their fixed position



Heated
(heat absorbed)

solid **liquid**

On increasing the temperature of solids , the kinetic energy of the particles increases . The particles start vibrating with greater speed . The energy supplied by heat over comes the forces of attraction between the particles . The particles leave their fixed positions and start moving more freely . A stage is reached when the solid melts and is converted to a liquid .

SUBLIMATION

- A process of solid changes directly to gas without becoming a liquid

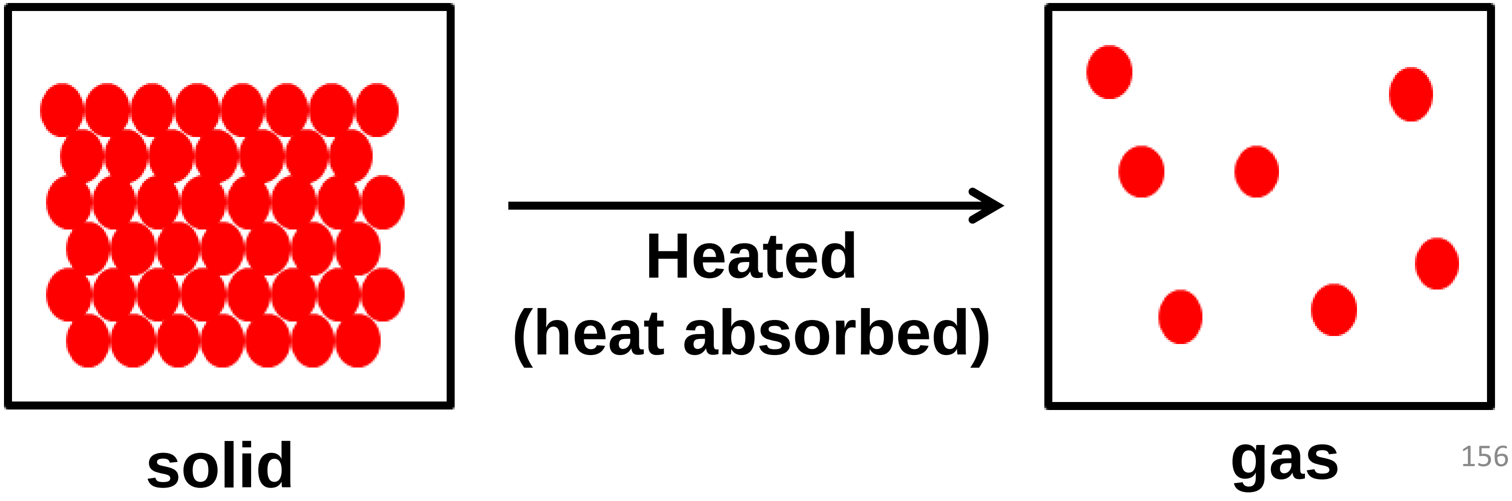
EXAMPLE:

$CO_2(s) \rightarrow CO_2(g)$

(dry ice)

- When a substance sublimes, molecules leave the surface of the solid, where they are held by less tightly than the interior

- The combination of intermolecular attractions and atmospheric pressure is not great enough to keep the particles near to one another when they leave solid state.



DEPOSITION

- A process of gas changes directly to solid without becoming a liquid

EXAMPLE: $I_2(g) \rightarrow I_2(s)$

When iodine vapor comes in contact with a cold surface, it deposits iodine crystals
$I_2(g) \rightarrow I_2(s)$

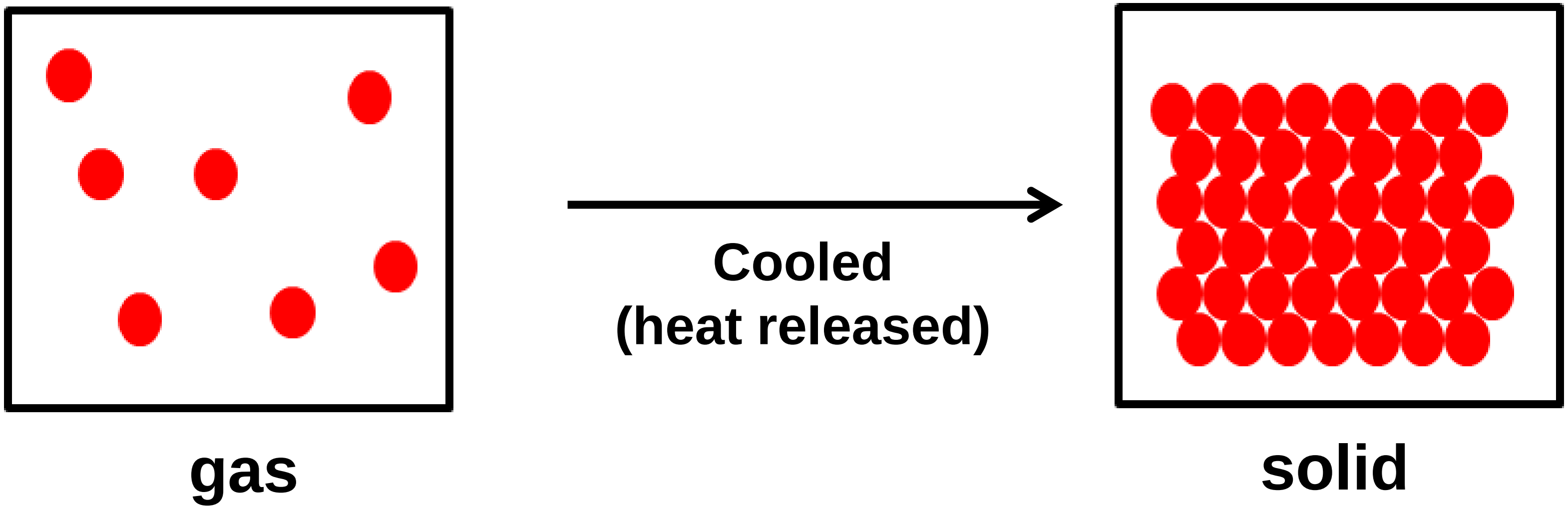
At ordinary atmospheric pressure, solid iodine sublimes: $I_2(s) \rightarrow I_2(g)$

- **When temperature of a gas is lowered, the kinetic energy of the gas molecules decreases.**

- **The kinetic energy unable to overcome the intermolecular forces between the molecules.**

- **The intermolecular forces are strong enough to hold the molecules in fixed position.**

- **The gas molecules change to solid.**



Cooled
(heat released)

gas

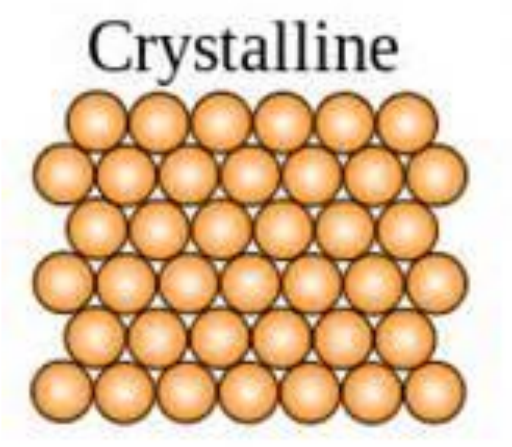
solid

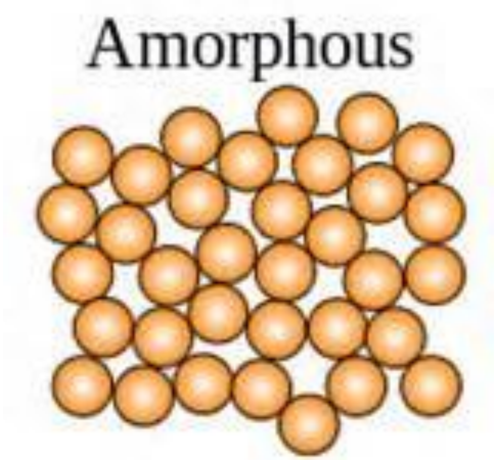
Thermodynamic process of phase transition in which gas transforms into solid without passing through the liquid phase. One example of deposition is the process by which, in sub freezing air, water vapor changes directly to ice without first becoming a liquid. This is how snow forms in clouds, as well as how frost form on the ground or other surfaces.

TYPES OF SOLID

Types of Solid

- **Crystalline Solid**
- **Amorphous Solid**





Differentiate Between Amorphous & Crystalline solid

	Crystalline Solid	**Amorphous Solid**
Particles arrangement	Atoms, ions or molecules are orderly arrangement.	Atoms, ions or molecules have no orderly structure/ random arrangement
Shape	Well-defined shape	Poorly defined shape
Examples	Ice, sugar, salt, diamond	Glass, charcoal, rubber
Types	1. Metallic solid 2. Ionic solid 3. Molecular covalent solid 4. Giant covalent solid	

TYPES OF CRYSTALLINE SOLIDS

Defined by the type of particles in the crystal

☞ based on the forces hold the particles :

- Metallic solid (Na, Fe)
- Ionic solid (NaCl, CaF_2)
- Molecular covalent solid (O_2, H_2O)
- Giant covalent solid (diamond, SiO_2)

METALLIC SOLID

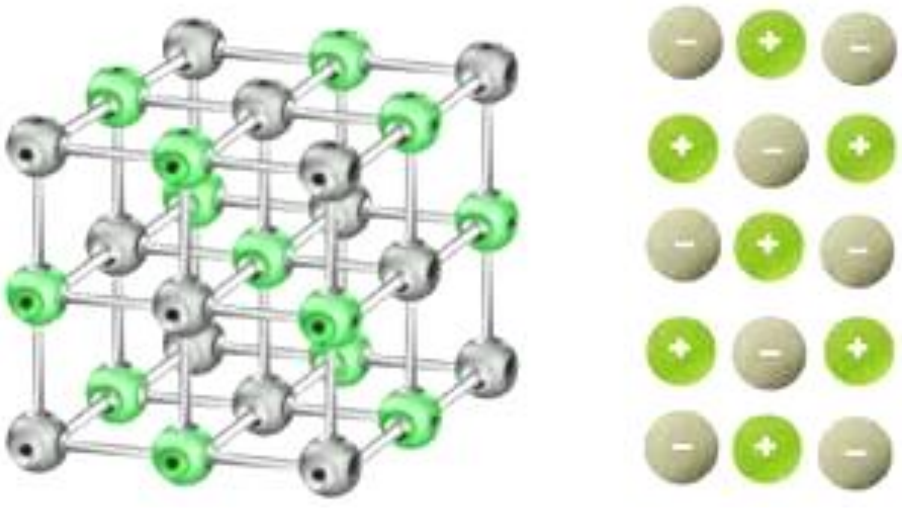
- Lattice points occupied by metal atoms
- Held together by strong metallic bond
- Soft to hard, low to very high melting point
- Good conductor of heat and electricity
- Malleable, ductile

EXAMPLE:

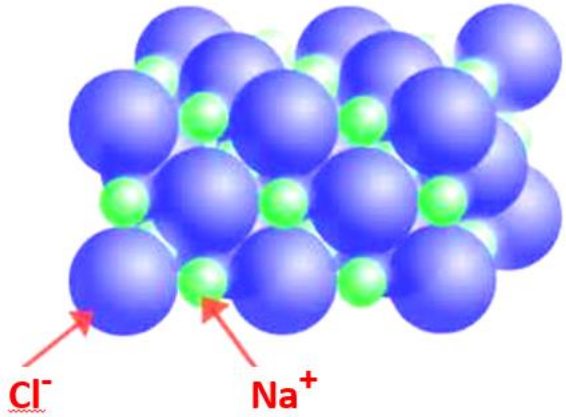
Na, Zn, Fe

IONIC SOLID

- Lattice points occupied by cations and anions
- Held together by ionic bonds (strong electrostatic attraction between cations and anions)
- Hard, brittle, high melting point
- Poor conductor of heat and electricity (in solid state) but does so in molten state or in aqueous solution due to the free moving ions



Lattice points in an ionic solid



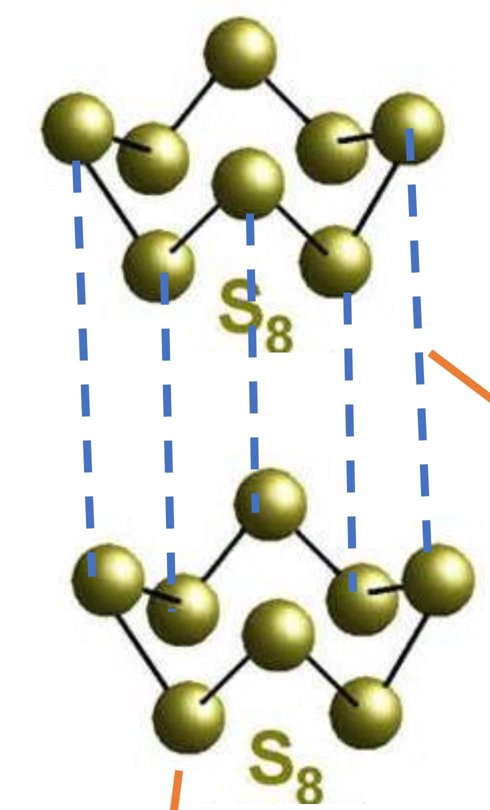
Example of ionic solid : NaCl

MOLECULAR COVALENT SOLID

- Lattice points occupied by molecules
- Held together by intermolecular forces
  (by weak van der Waals forces or hydrogen bond
- Fairly soft, low to moderate melting point
- Poor conductor of heat and electricity

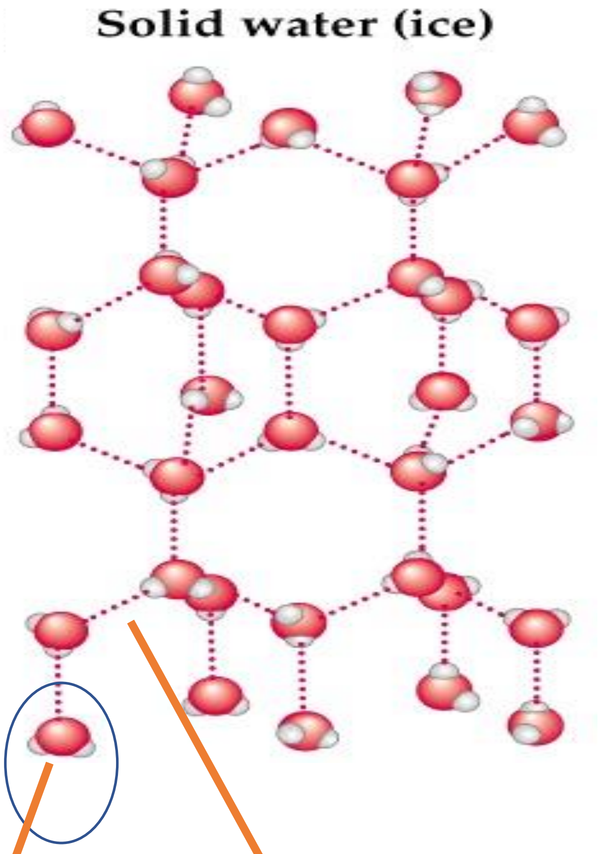
EXAMPLE: Nonpolar molecule : P_4, S_8

Polar molecule : H_2O (ice)



Van der Waals forces acting between molecules of S_8

Sulphur atom, S



Water molecule, H_2O

Hydrogen bond acting between water molecules in solid H_2O (ice)

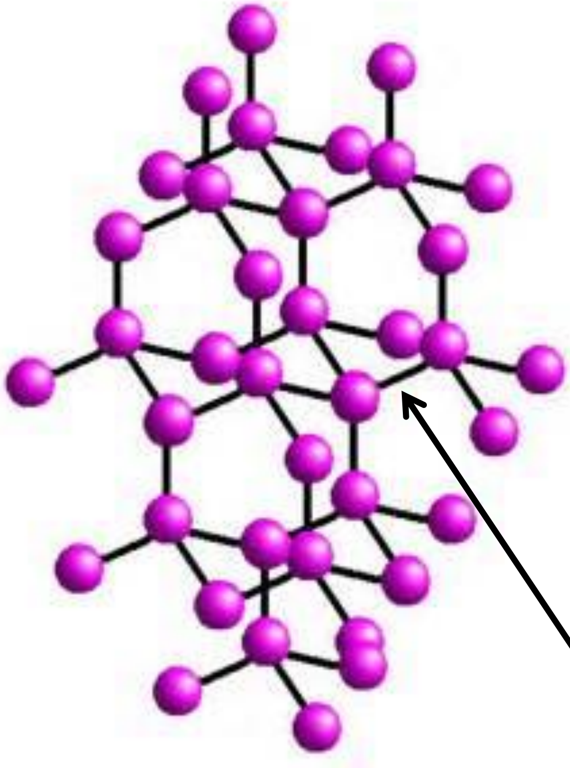
GIANT COVALENT SOLID (NETWORK SOLID)

- Also called network solid
- The particles are held together by strong covalent bonds and forms continuous network
- Very hard, very high melting point
- Usually poor conductor of heat and electricity

EXAMPLE:

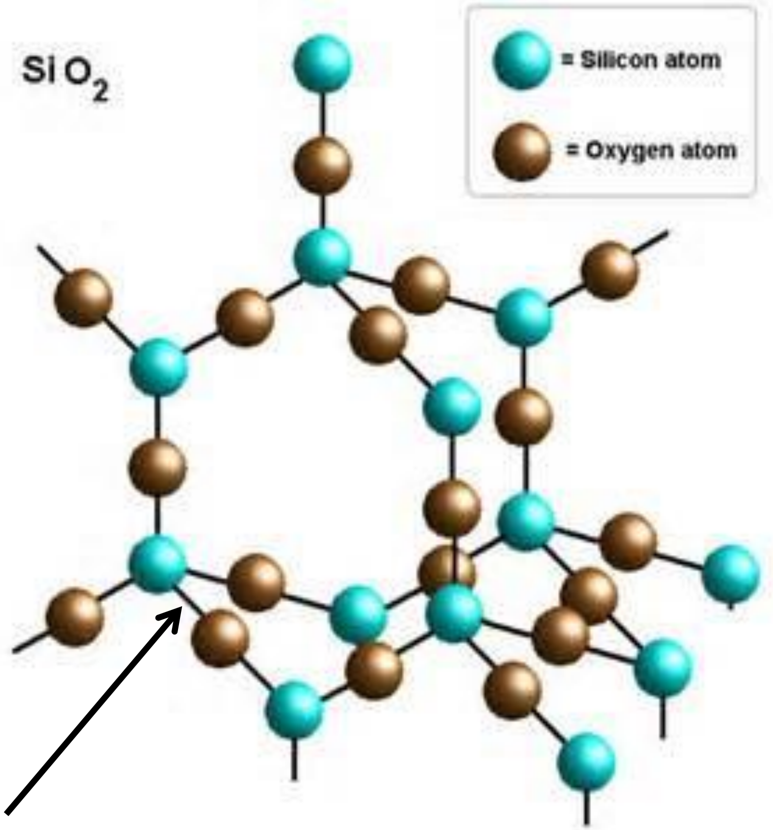
C allotrope (diamond)

SiO_2



C allotrope (diamond)

Covalent bond



SiO_2

5.4 Phase diagram

Phase Diagram

- **Define**
  - Phase
  - Triple point
  - Critical point
- **Phase Diagram**
  - Positive Slope – CO_2
  - Negative Slope – H_2O
- **Triple Point & Critical Point on the phase diagram**
- **Anomalous behaviour of H_2O**
- **Changes in Phase**
  - Temperature(constant P) @ ISOBARIC
  - Pressure (constant T) @ ISOTHERMAL

5.4 PHASE DIAGRAM

Teaching and learning outcomes

At the end of the lesson, student should be able to

5.4	Phase Diagram
a)	Define: (C1) i) phase ii) triple point iii) critical point
b)	Identify triple point and critical point on the phase diagram. (C2)
c)	Sketch the phase diagram of H_2O and $CO_{2.}$ (C3)
d)	Compare the phase diagram of H_2O with $CO_{2.}$ (C4)
e)	Explain the anomalous behaviour of H_2O. (C2, C3)
f)	Describe the changes in phase with respect to: (C1, C2) i. Temperature (at constant pressure) ii. Pressure (at constant temperature)

PHASE

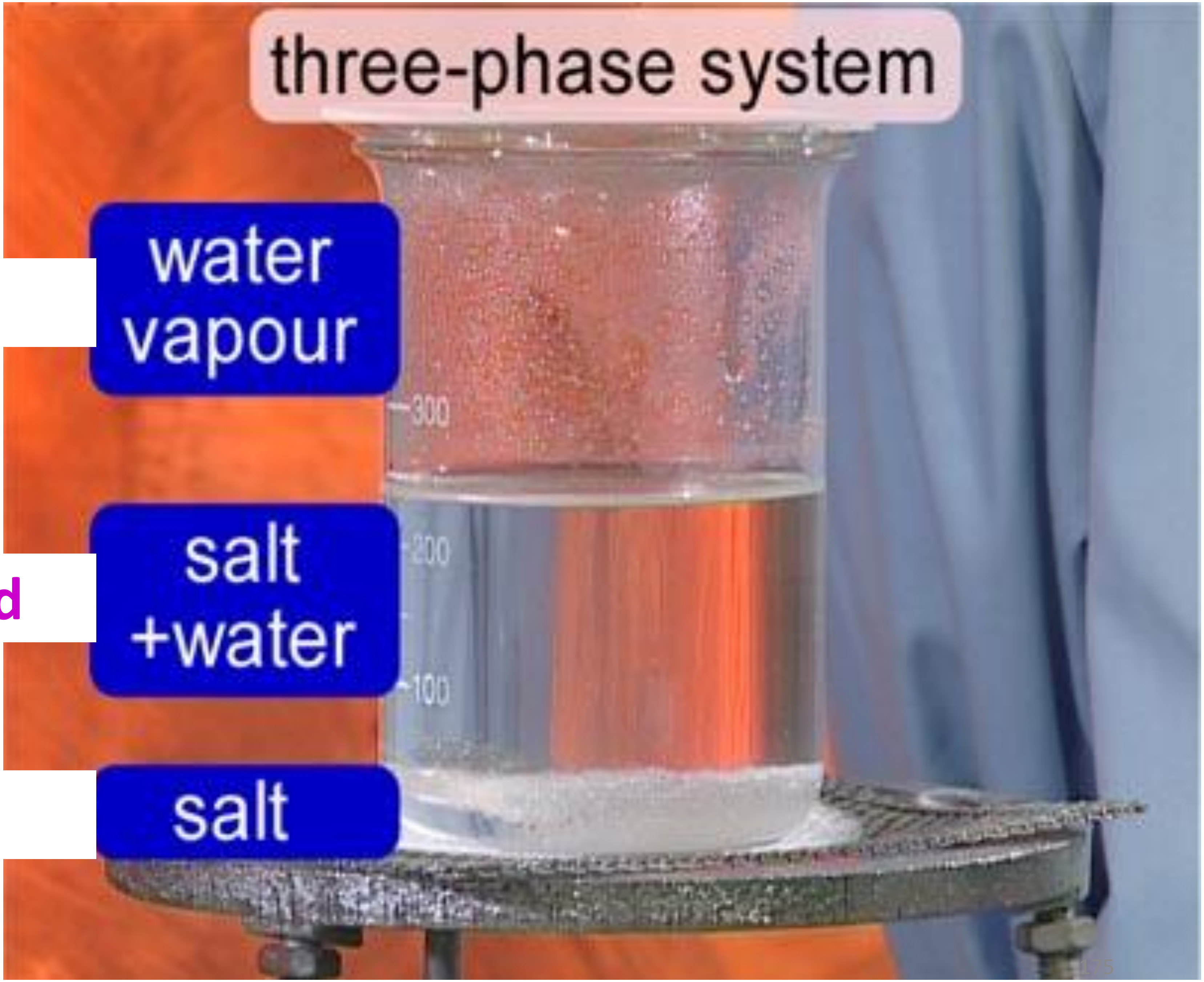
- A homogenous part of a system in contact with other parts of the system but separated from by well-defined boundary
- A phase is homogeneous part
- Phase consist of :
  - Solid
  - Liquid
  - Gas

EXAMPLE :

two phases

Solid phase : ice

Liquid phase : water



gas

liquid

solid

How many phase?

one phase (liquid)

Oil + water

two phase (liquid)

(but heterogenous)

How many phase?

One gas only

one phase (gas)

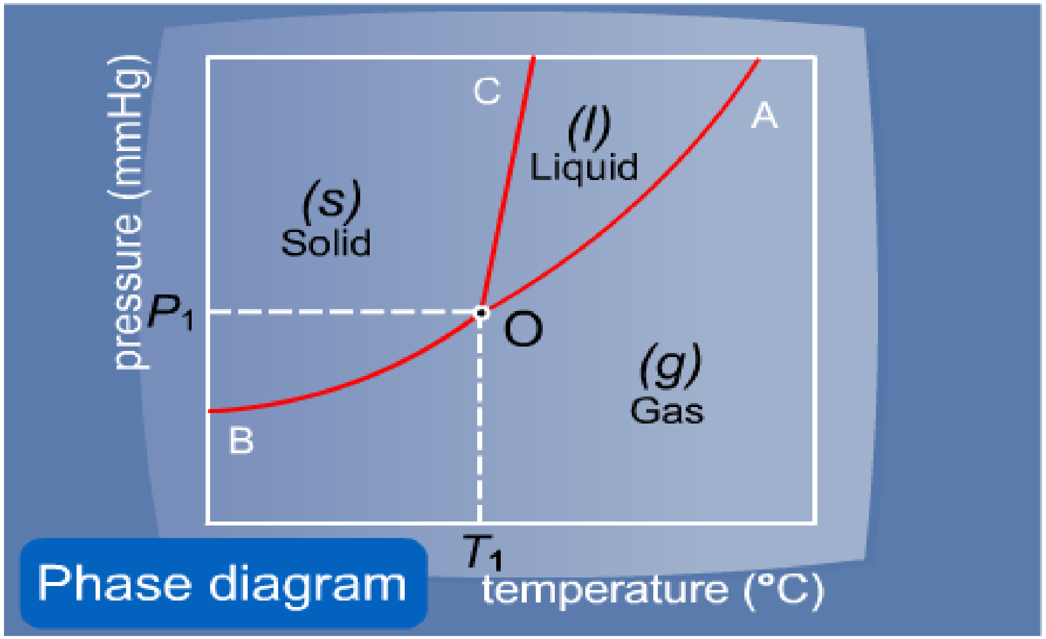
Mixture of gases

one phase (gas)

System	**Phase**	**Description**
Mixture of gases	1	Gases are well mixed. We cannot see the boundary between the gases.
Two liquids Do not mix (oil-water)	2	We can see the boundary between the two liquids.
Two liquids well mixed (alcohol-water)	1	We cannot see the boundary of the solution.

PHASE DIAGRAM

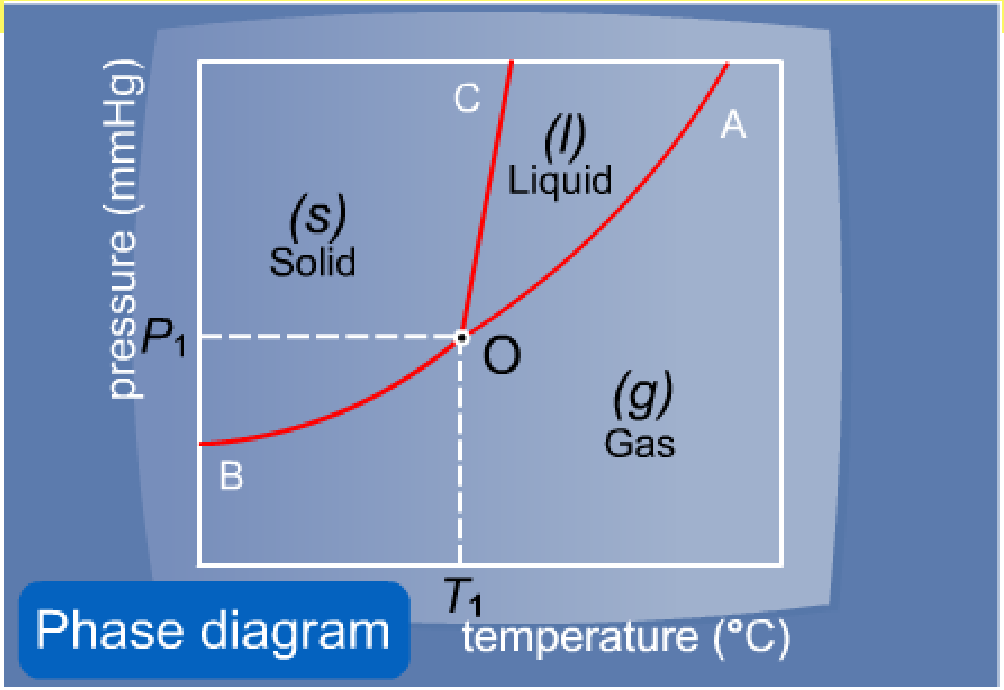
- A diagram showing the **regions** of **pressure** and **temperature** at which a substance exists as a solid, liquid or vapor phases are **thermodynamically stable.**



Phase diagram

It combines equilibrium curves of:

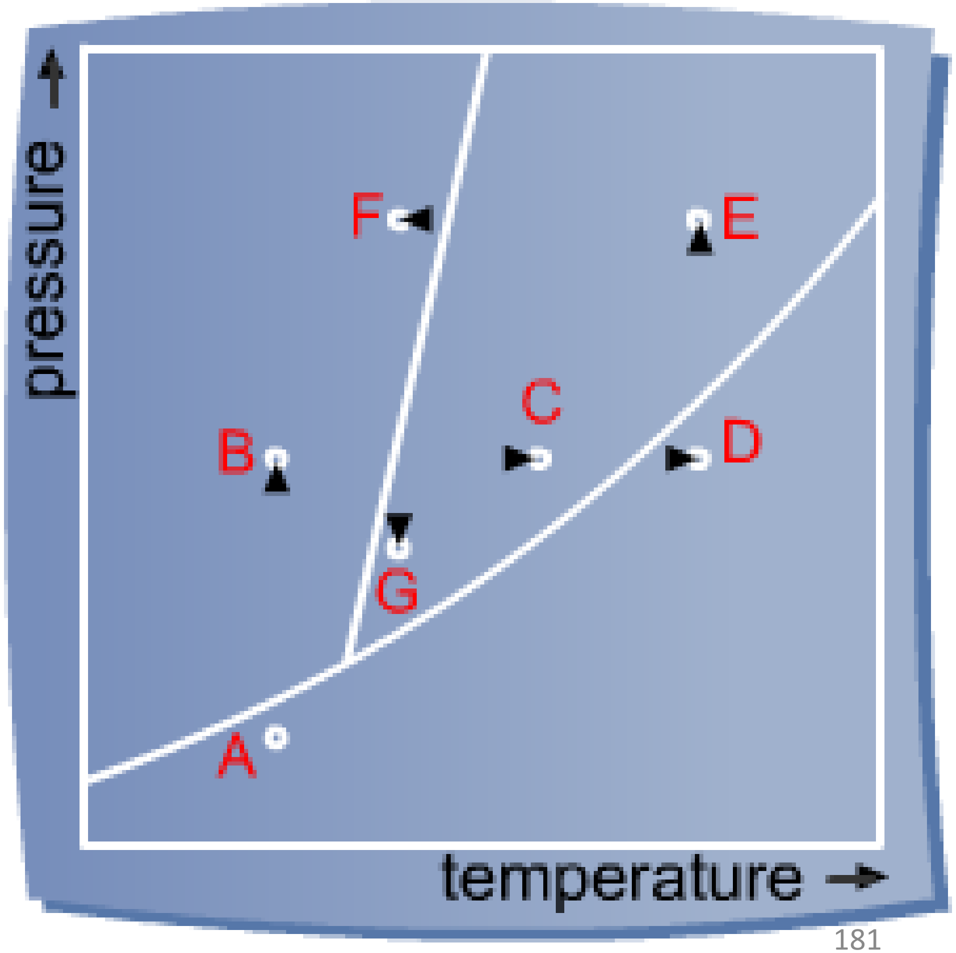
- liquid $\rightleftharpoons$ gas : Line OA
- solid $\rightleftharpoons$ liquid: Line OC
- solid $\rightleftharpoons$ gas : Line OB



EXAMPLE 1 :

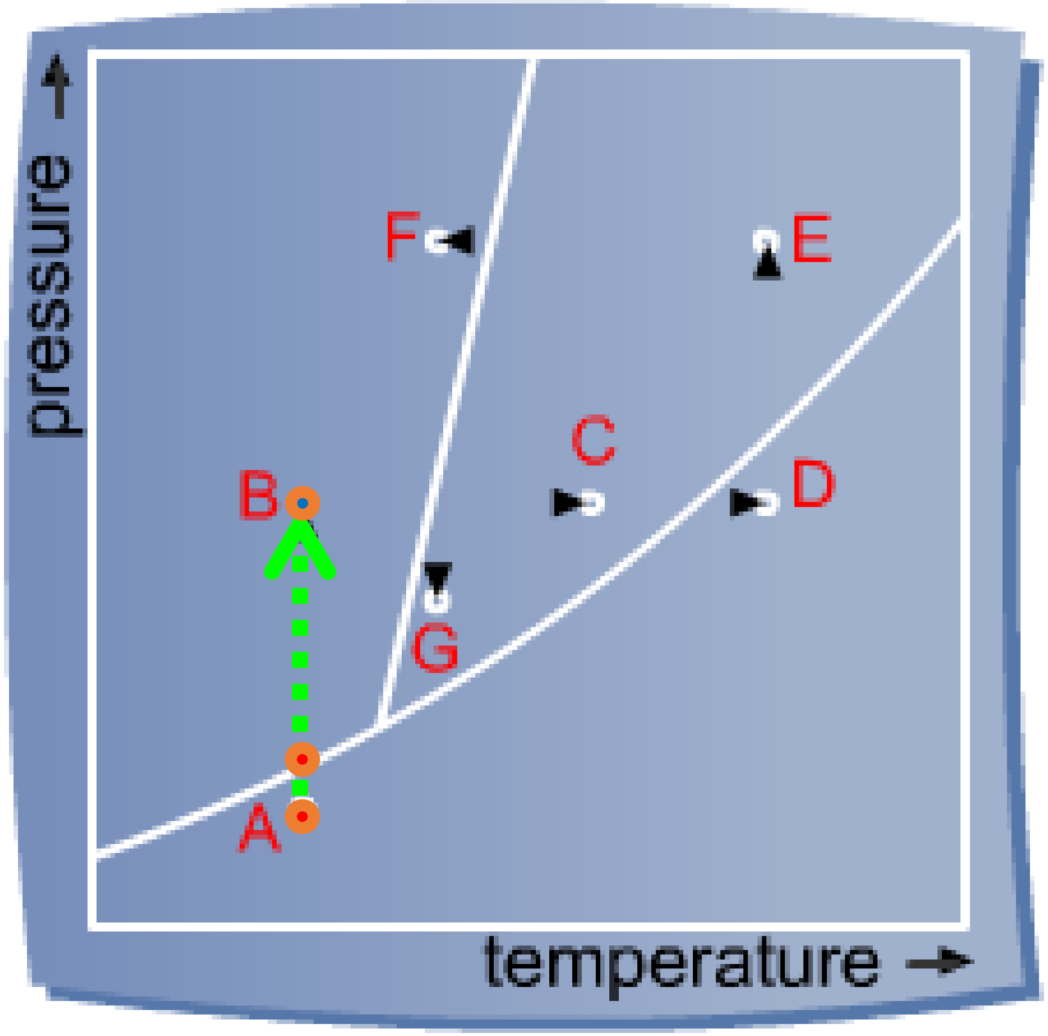
Predict the phase changes.

(a) A-B
(b) B-C
(c) C-D
(d) D-E
(e) E-F
(f) F-G



ANSWER 1:

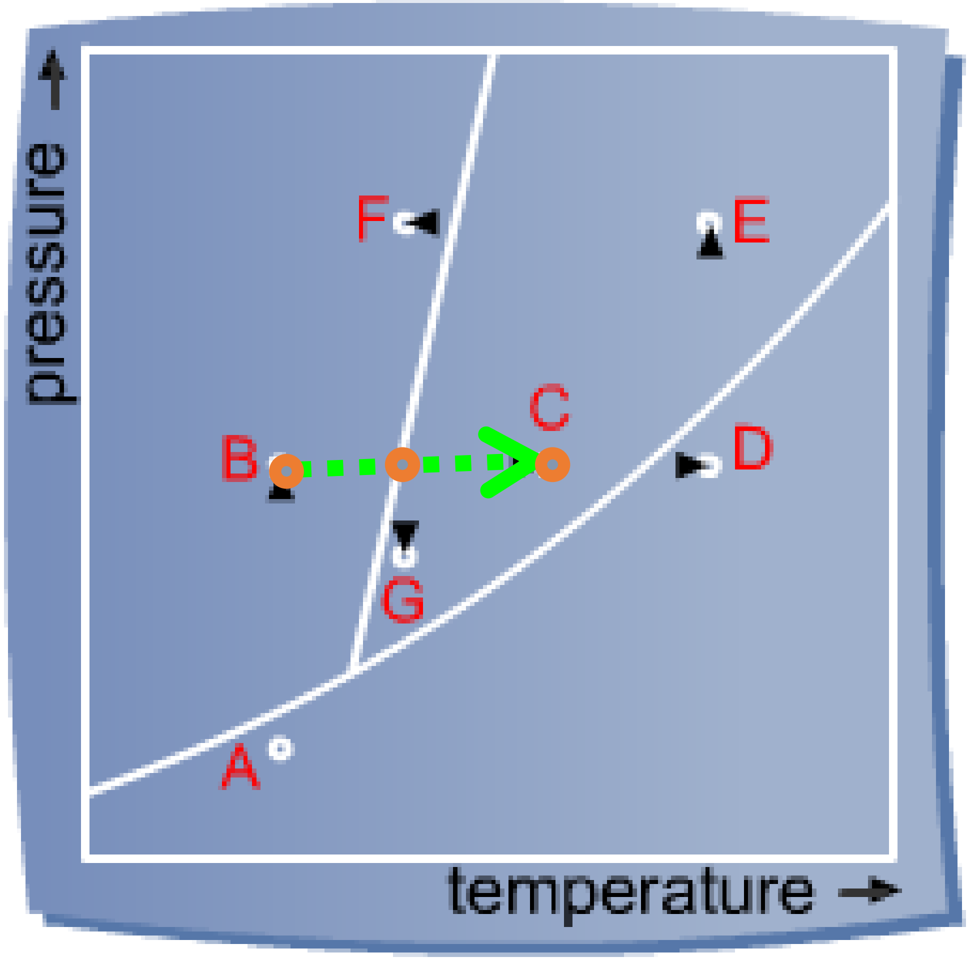
(a) A-B



☞ (gas) → (gas ⇌ solid) → (solid)

ANSWER 1:

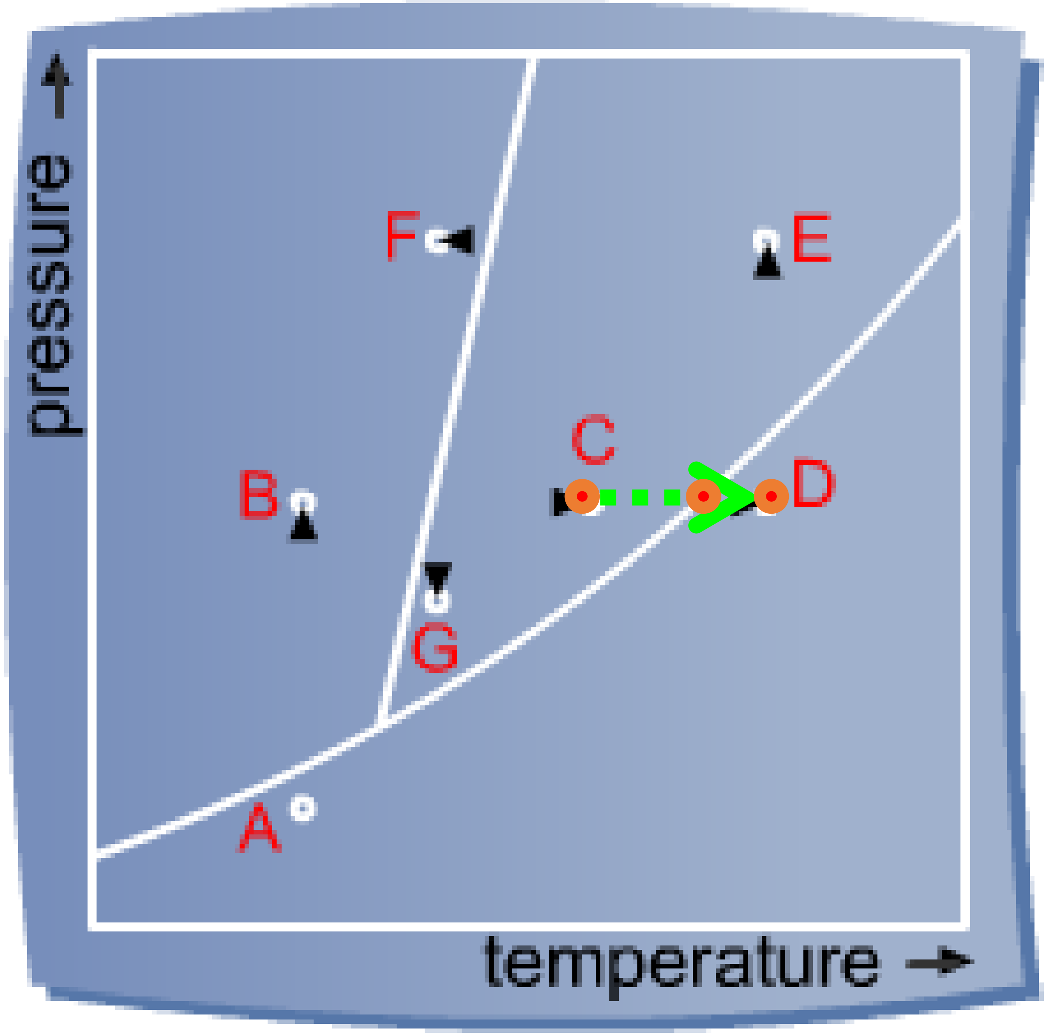
(b) B-C



☞ (solid) → (solid ⇌ liquid) → (liquid)

ANSWER 1:

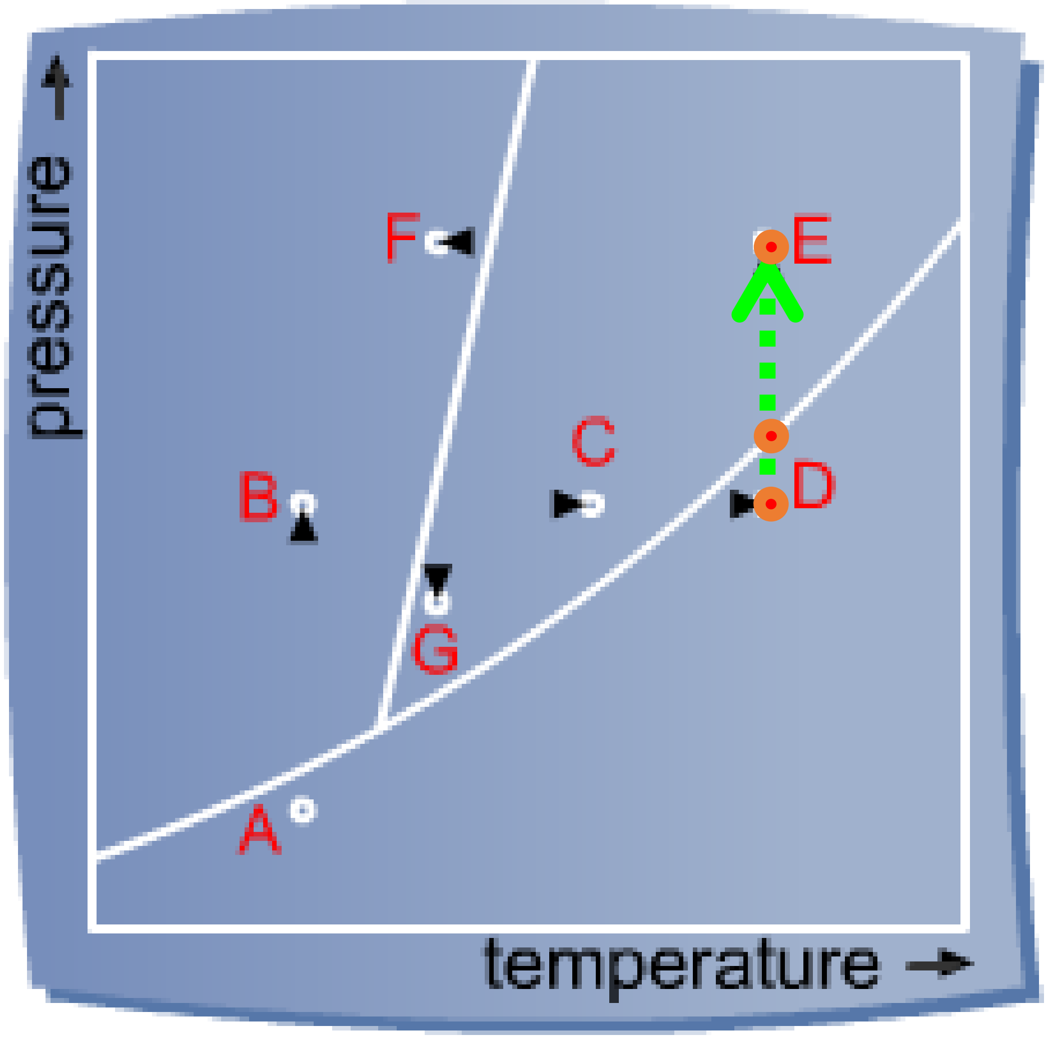
(c) C-D



☞ (liquid) → (liquid ⇌ gas) → (gas)

ANSWER 1:

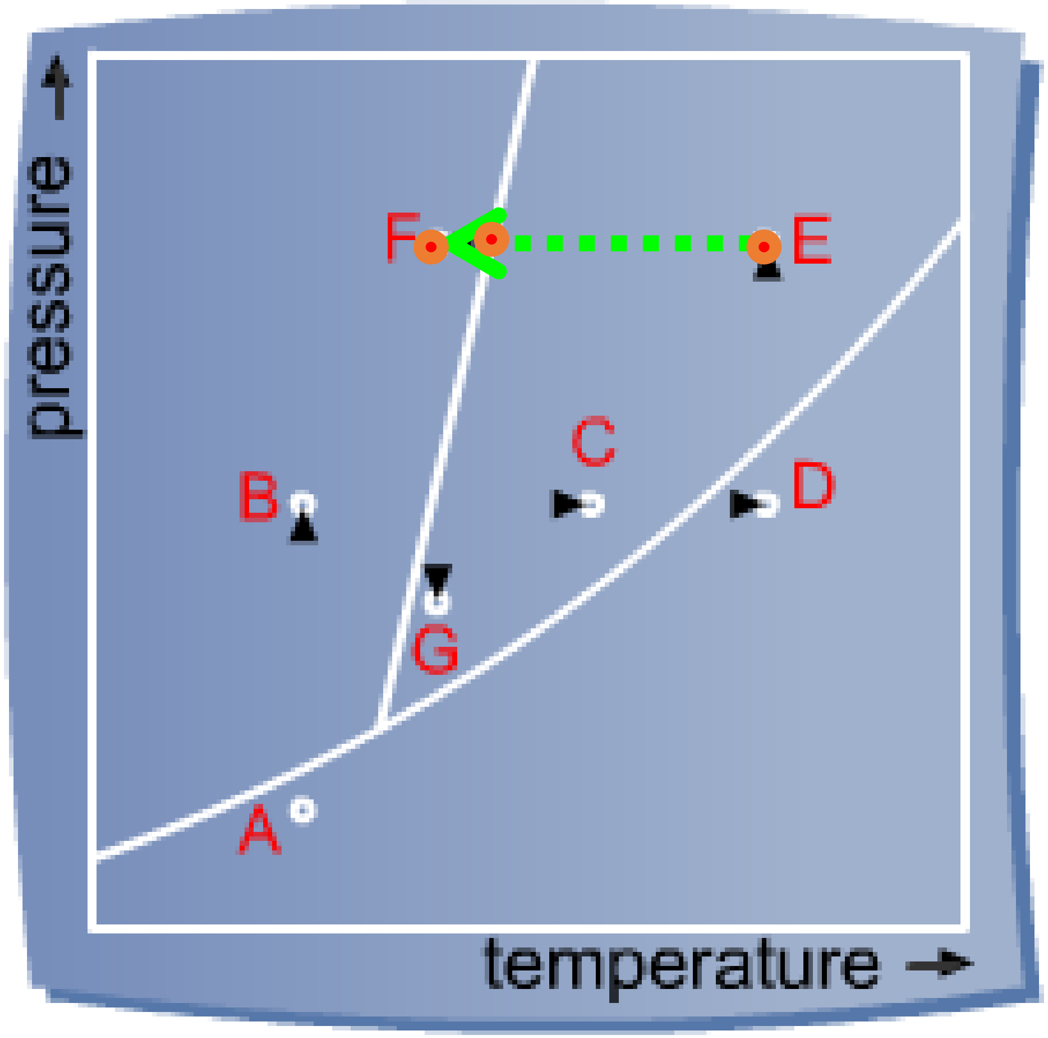
(d) D-E



☞ (gas) → (gas ⇌ liquid) → (liquid)

ANSWER 1:

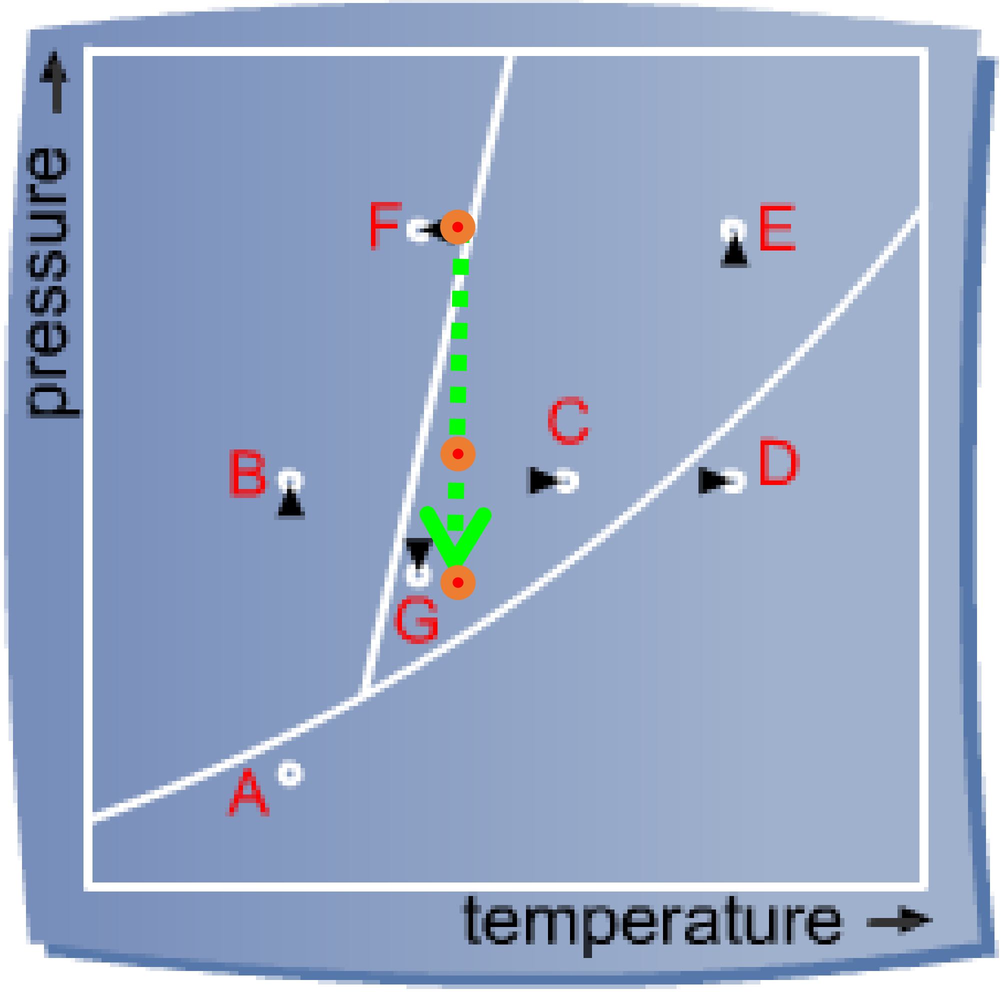
(e) E-F



☞ (liquid) → (liquid ⇌ solid) → (solid)

ANSWER 1:

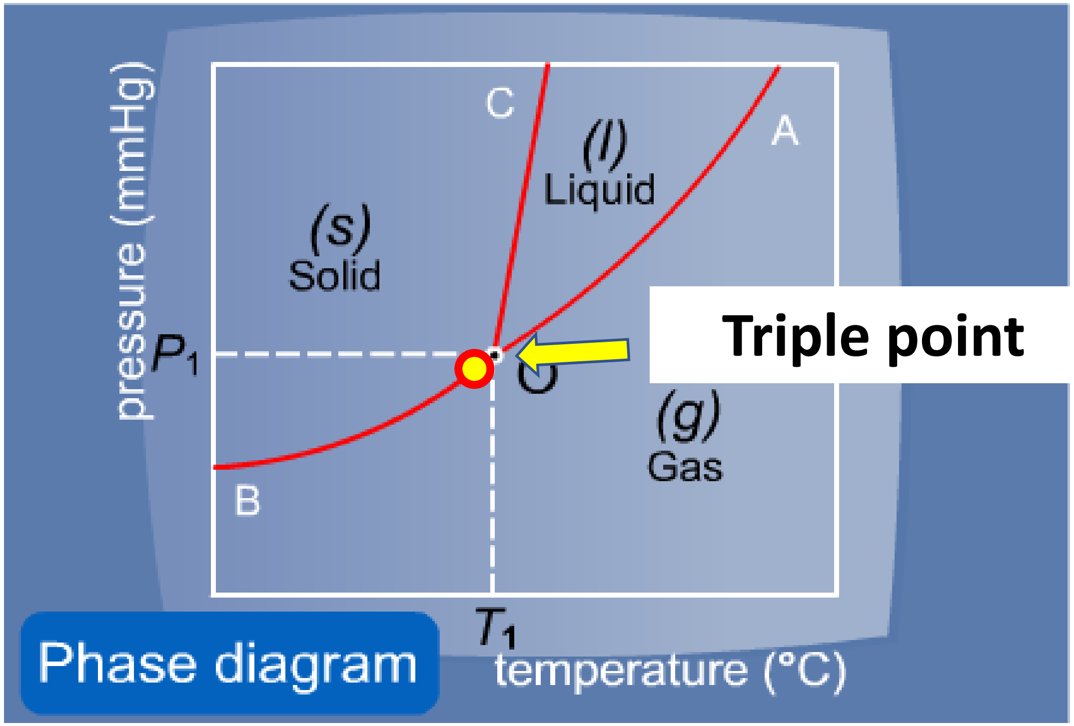
(f) F-G



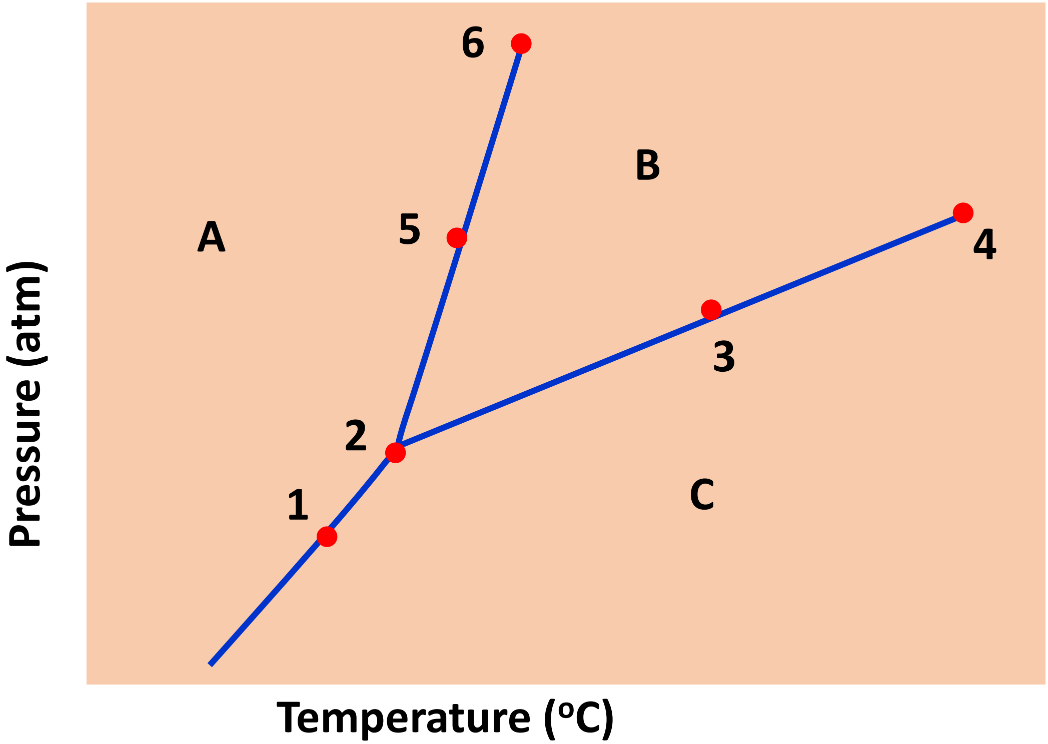
☞ (solid) → (solid ⇌ liquid) → (liquid)

TRIPLE POINT

- The temperature and pressure at which **solid**, **liquid** and **vapor** simultaneously exist in **equilibrium**



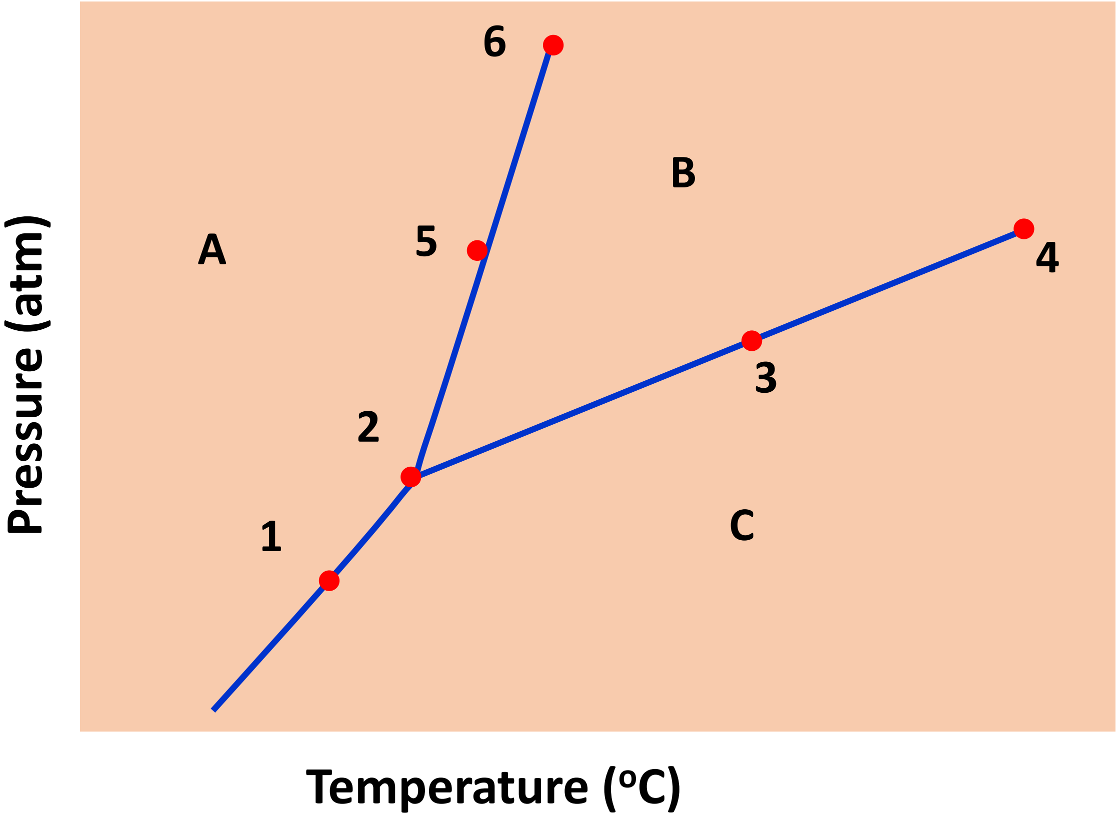
EXAMPLE 2 :



(a) What phase is present in region A? Region B? Region C?

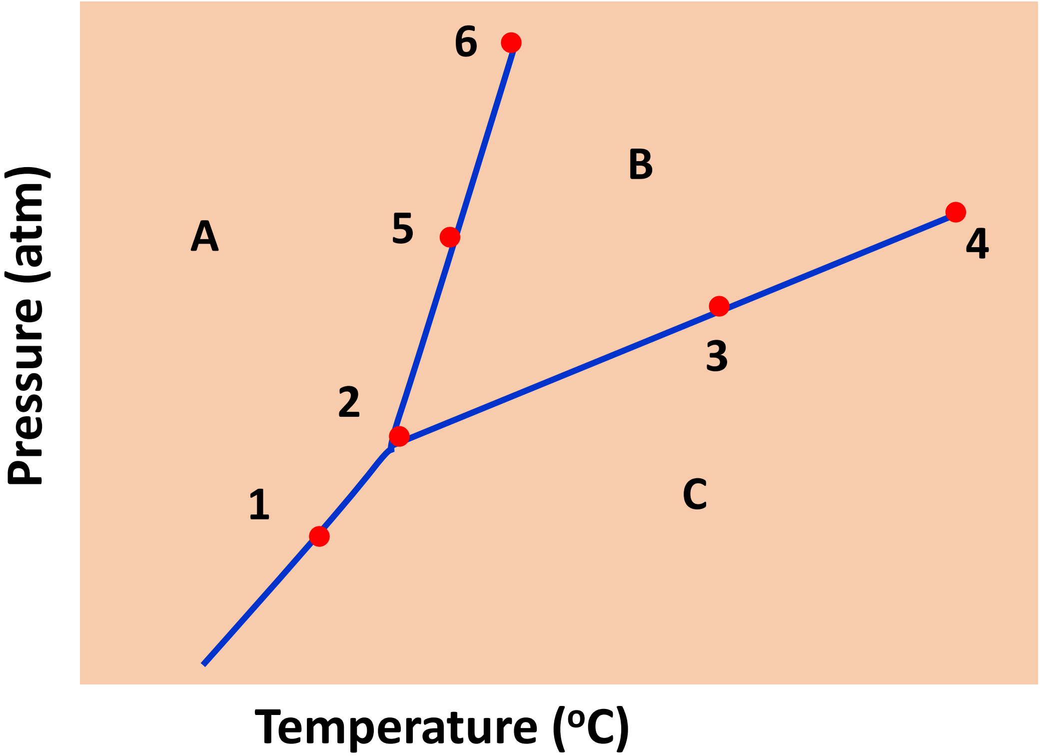
(b) What phases are in equilibrium at point 1? Point 2? Point 3? Point 5?

ANSWER 2:



(a) A: solid B: liquid C: gas

ANSWER 2:

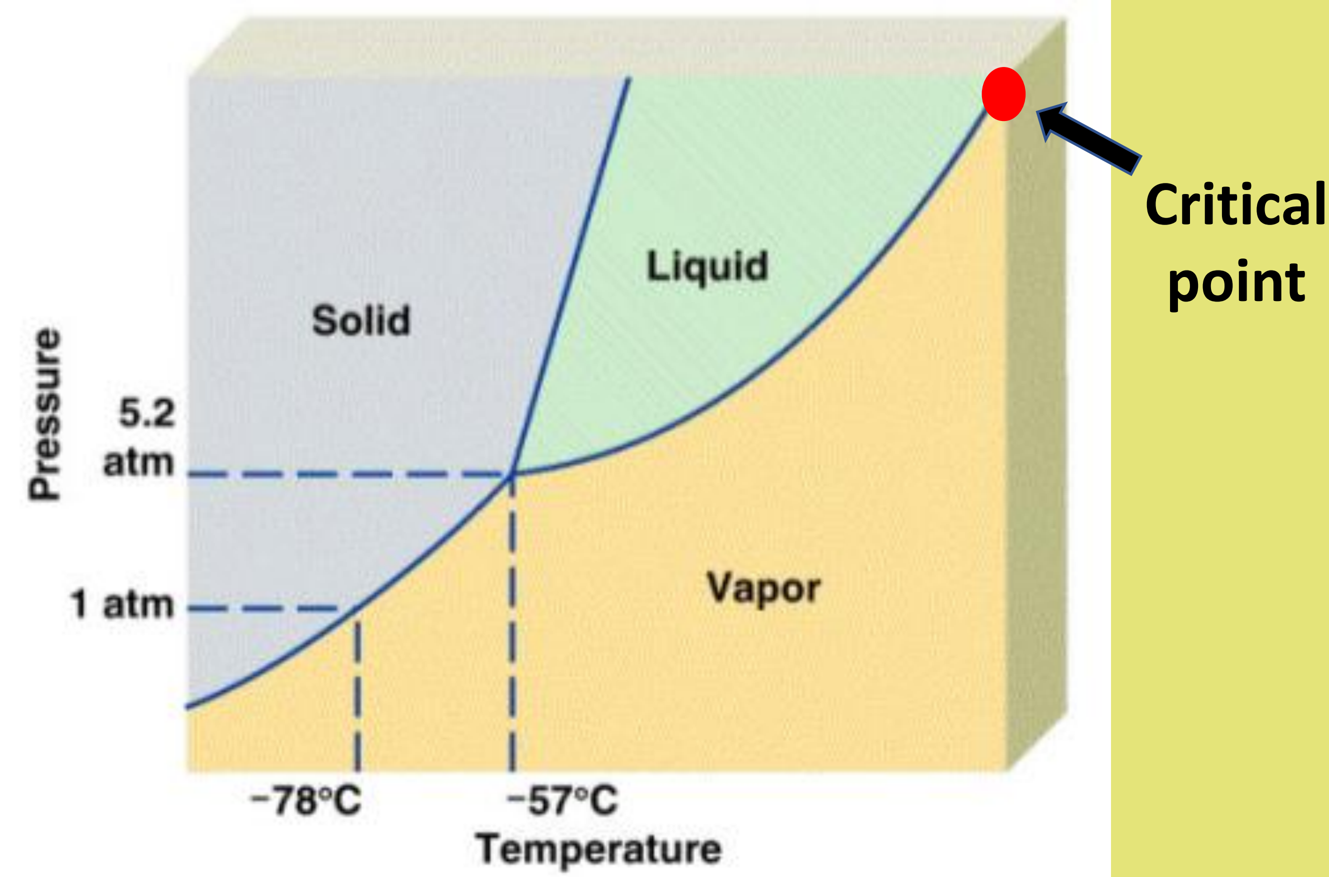


(b)

Point 1:	solid and gas
Point 2:	solid, liquid, and gas
Point 3:	liquid and gas
Point 5:	solid and liquid

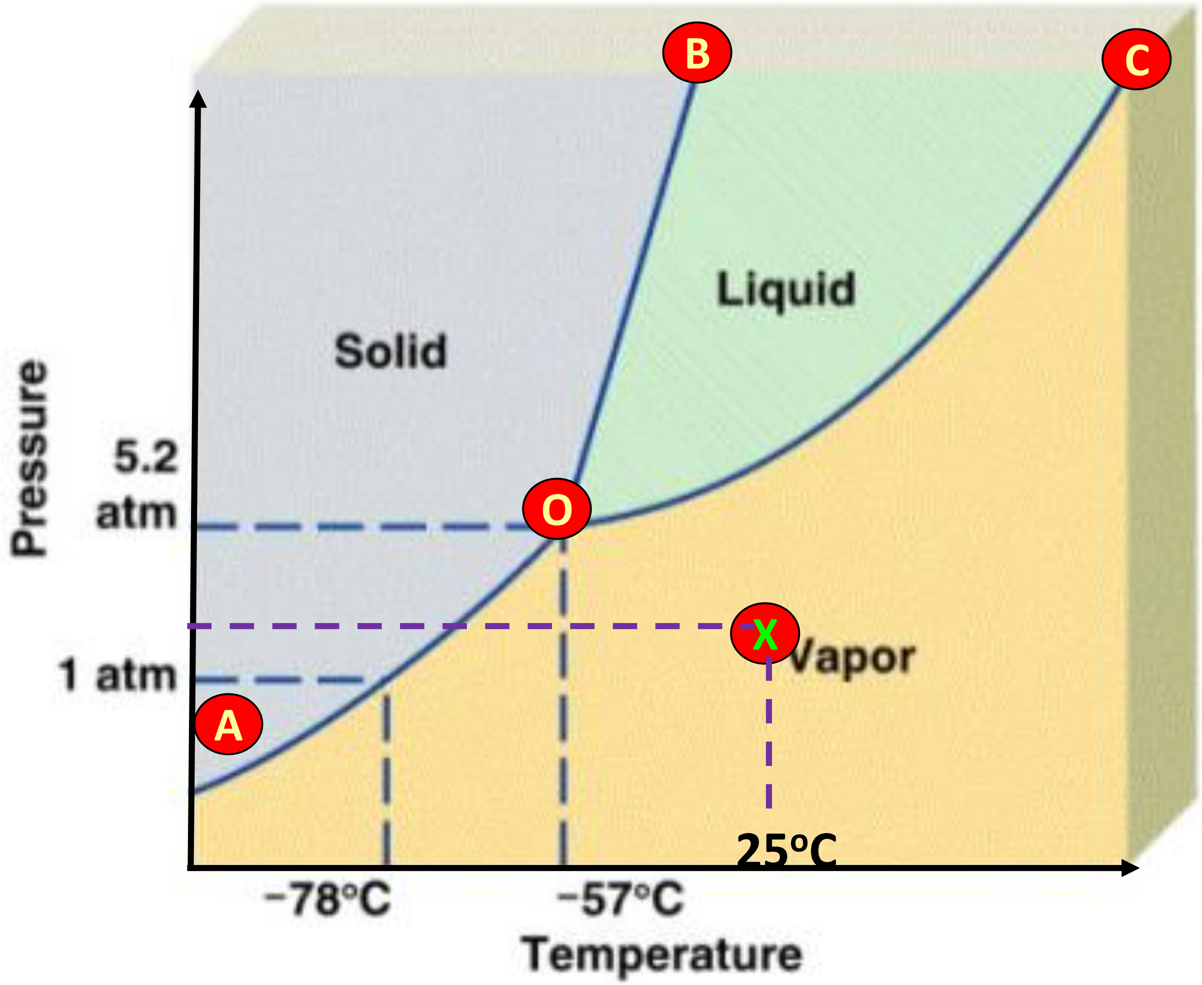
CRITICAL POINT

- The highest temperature and pressure at which a gas can be liquefied
- The end of liquid–gas line
- Density of liquid equal to density of gas
- Above the critical point, the liquid phase cannot exist, regardless the pressure
  - Phase boundaries disappears
  - Separate liquid and gas phase no longer exist

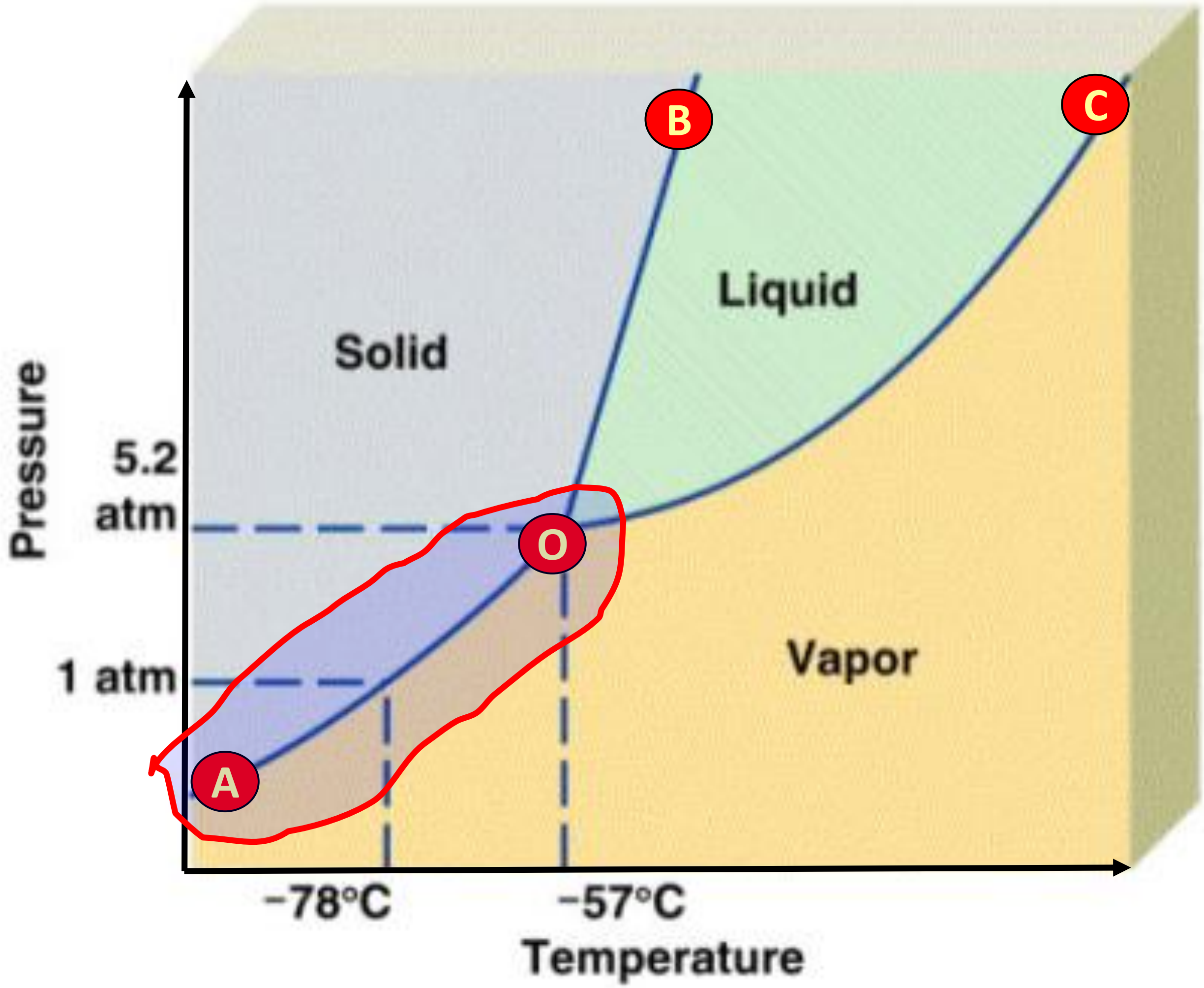


At this point, the average kinetic energy of the molecules is so high that the vapor cannot be condensed no matter how much pressure is applied

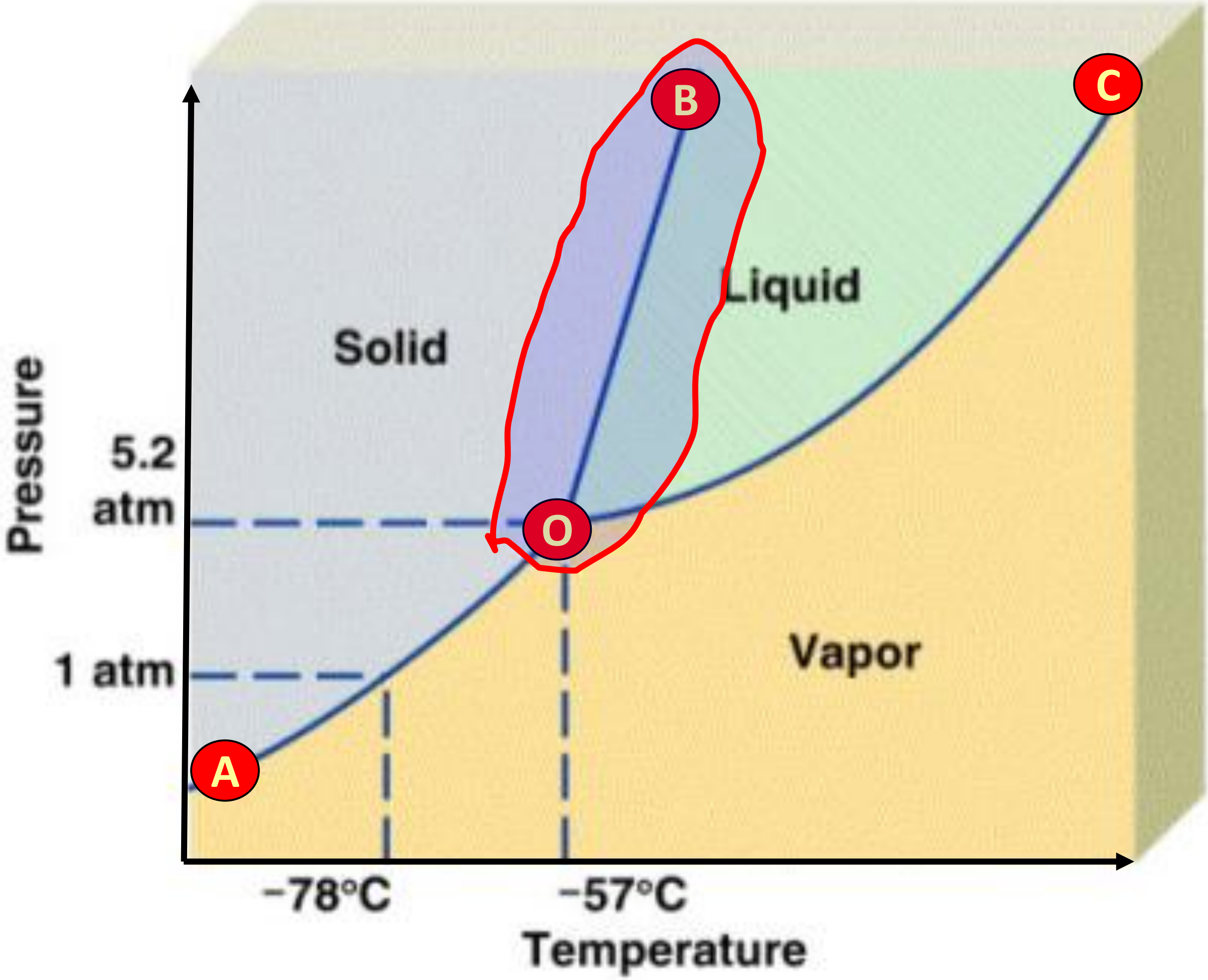
PHASE DIAGRAM OF CO_2



- CO_2 is a gas under normal conditions (25°C, 1 atm)

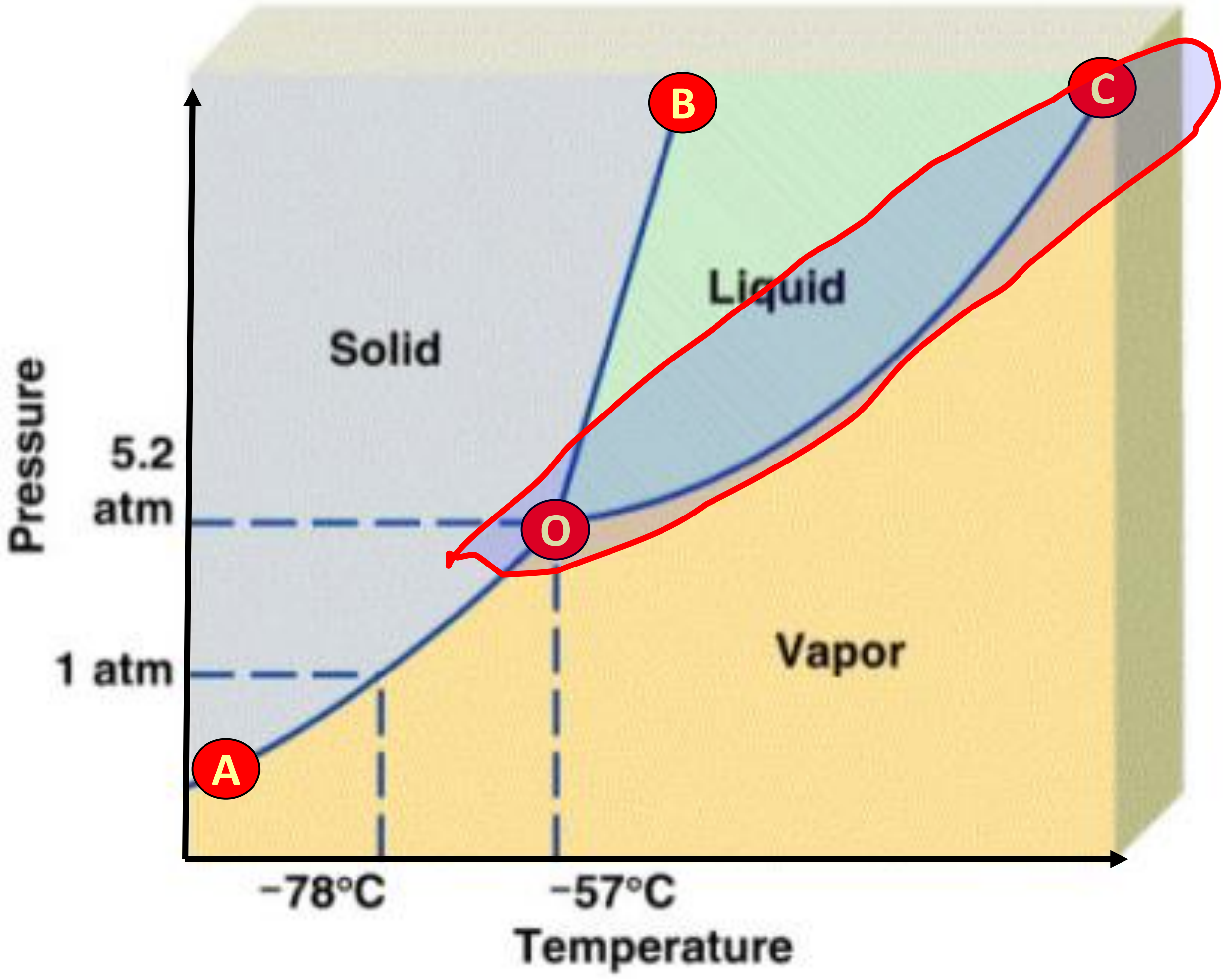


- Line OA: $CO_2(s) \rightleftharpoons CO_2(g)$
- Solid and gas exist in equilibrium



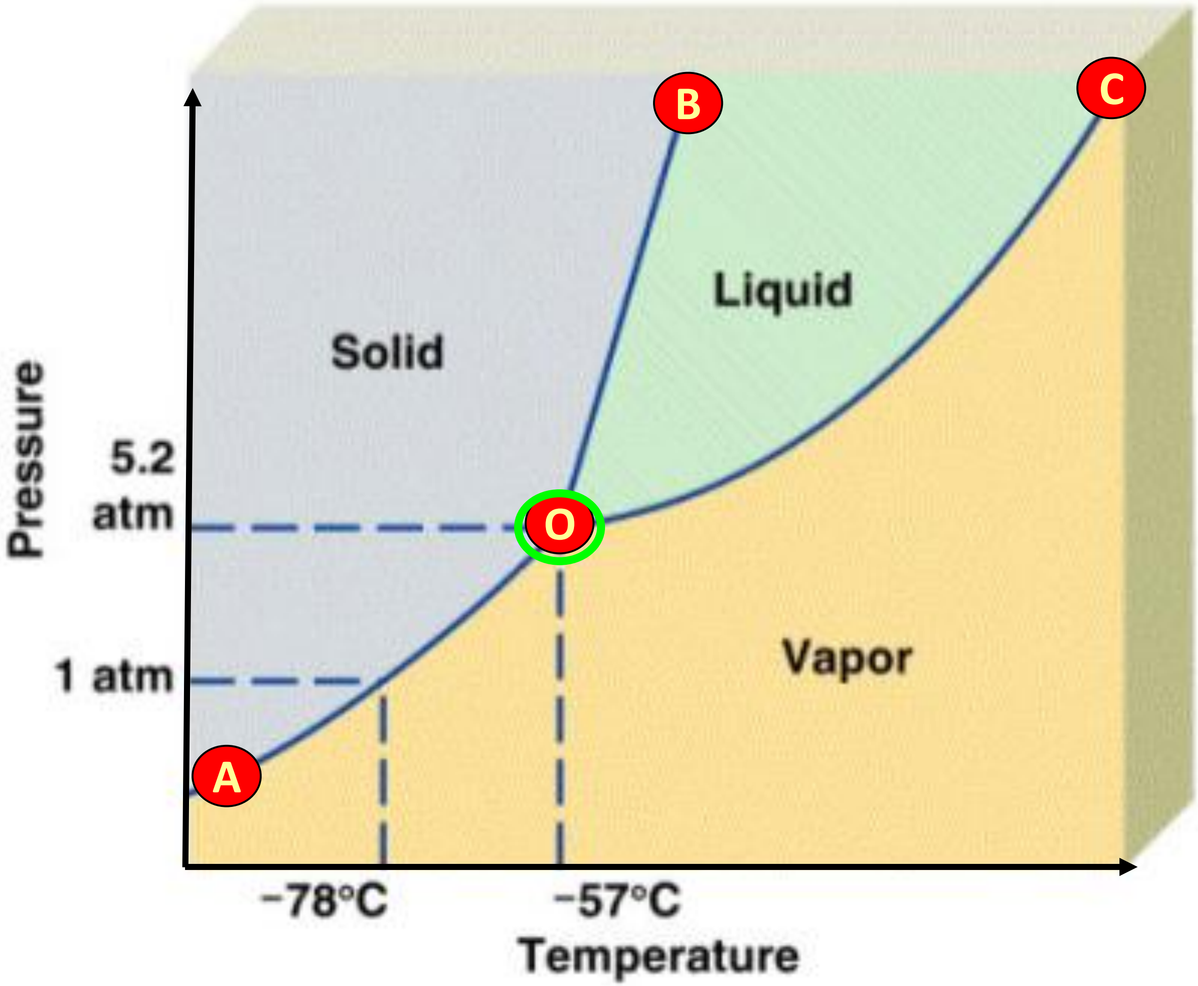
- Line OB: $CO_2(s) \rightleftharpoons CO_2(l)$ (melting / freezing line)

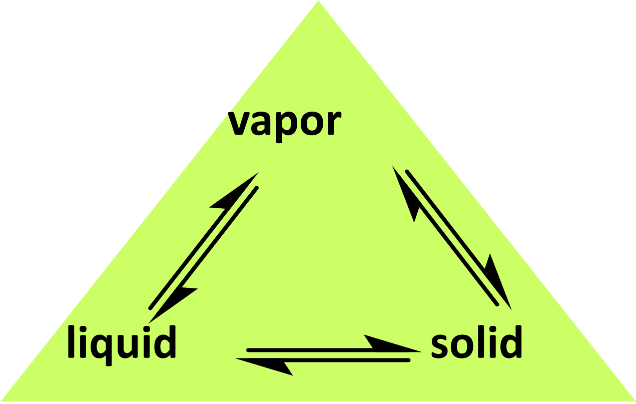
☞ solid and liquid exist in equilibrium



- Line OC: $CO_2(l) \rightleftharpoons CO_2(g)$ (boiling curve)

☞ liquid and gas exist in equilibrium

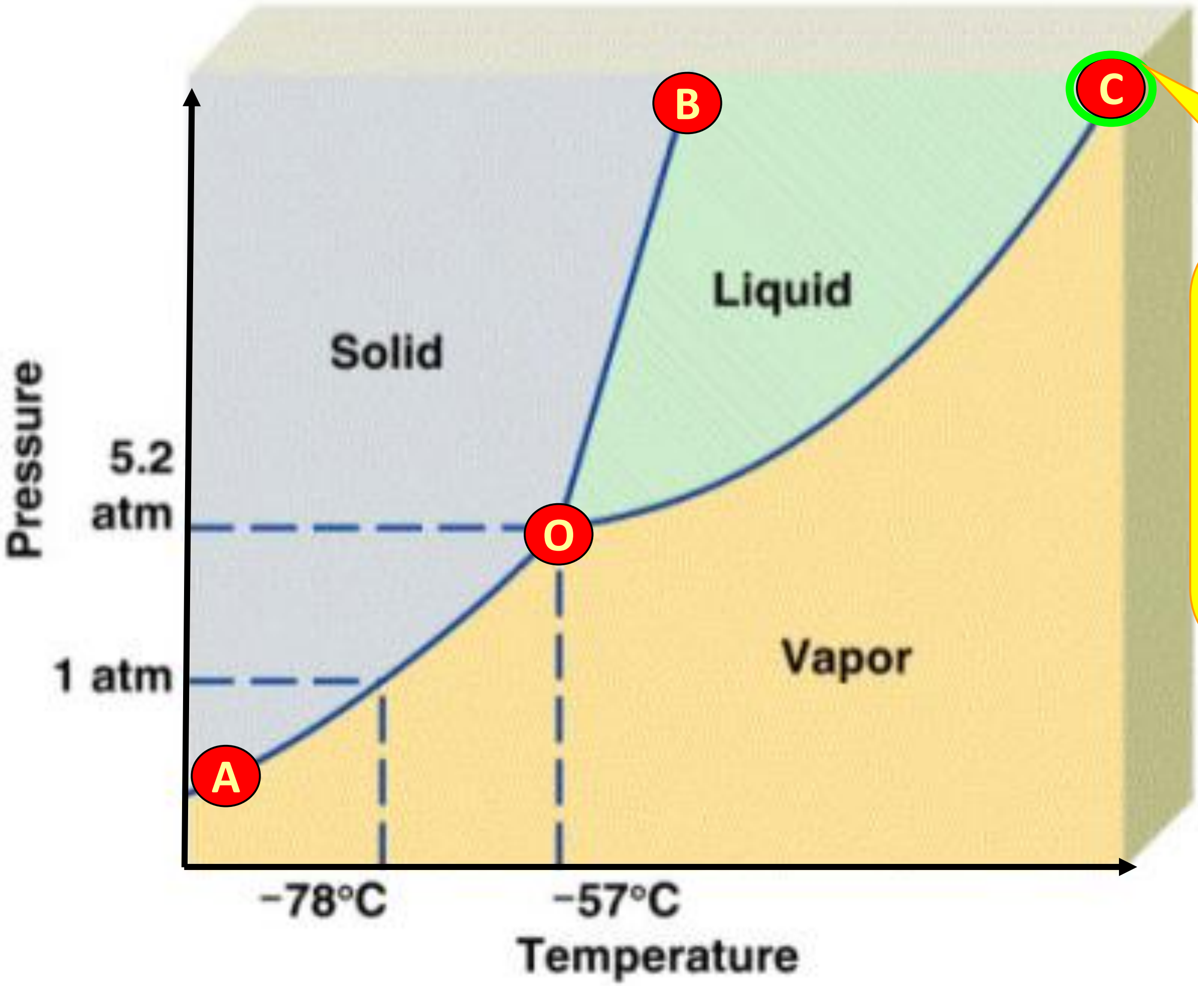



vapor

liquid

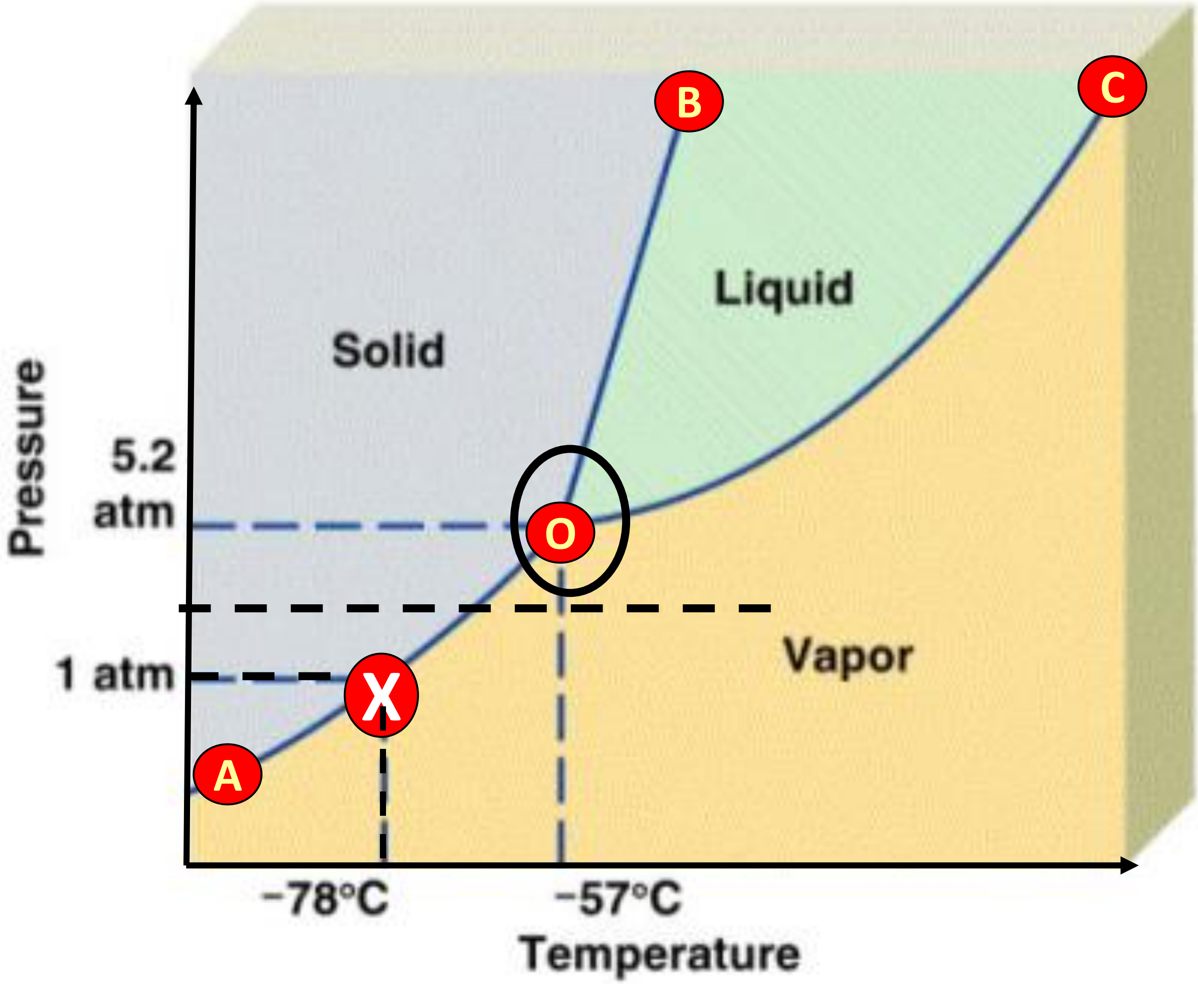
solid

- Point O = Triple Point (–56.4°C, 5.2 atm)
- Solid, liquid and vapor exist in equilibrium

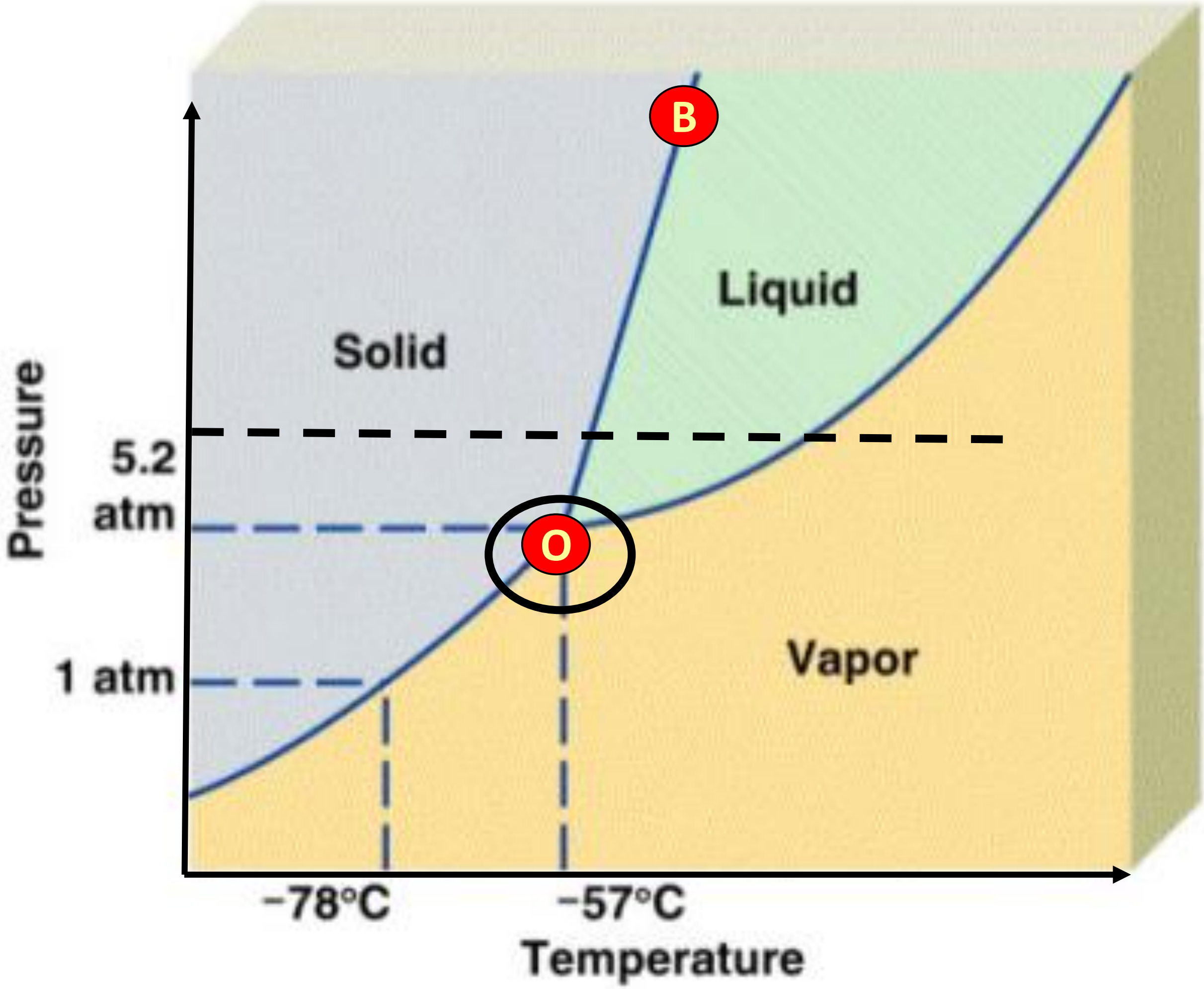


Above the critical point, the **liquid phase cannot exist**, regardless the pressure

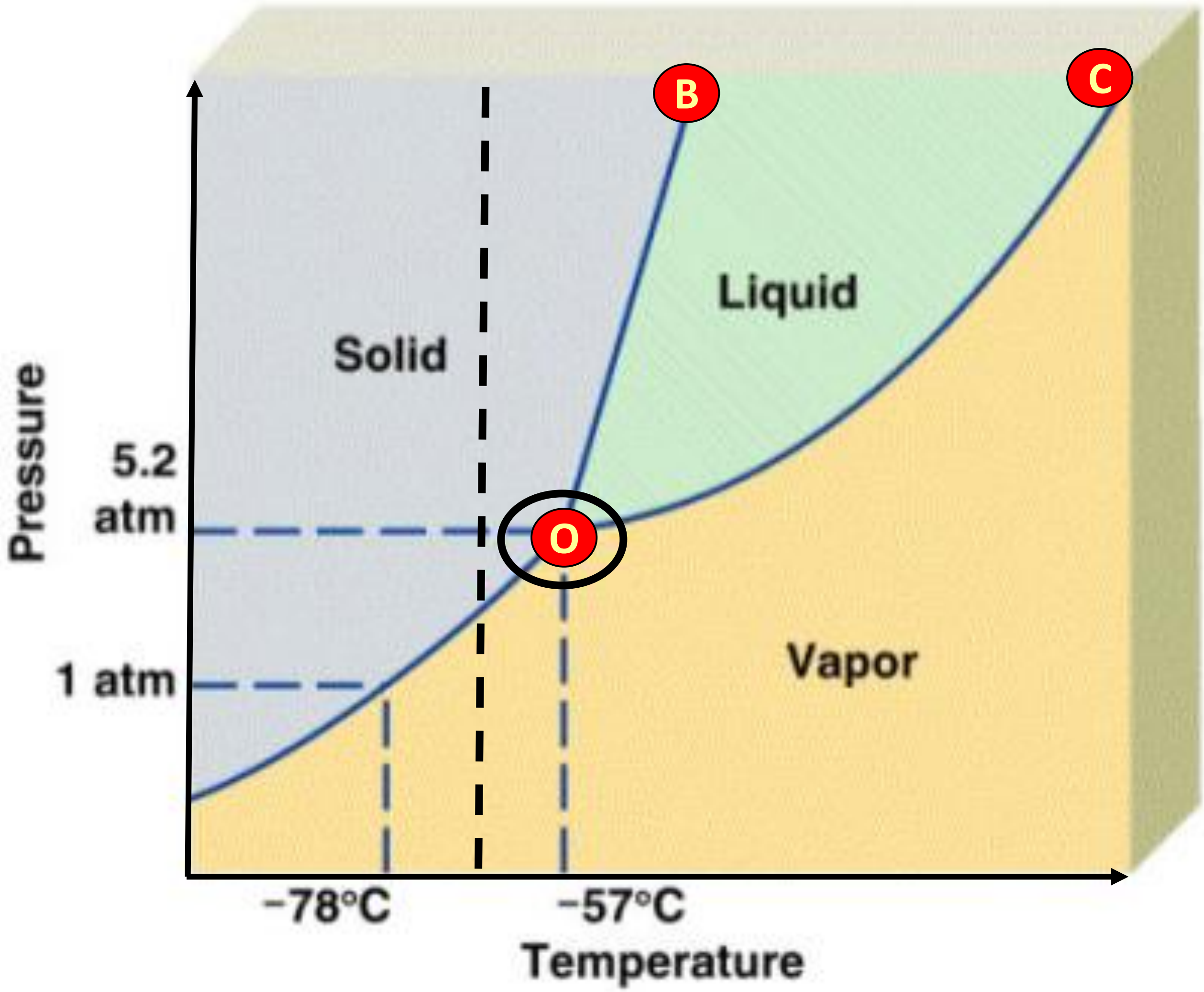
- **Point C: Critical Point (31°C, 73 atm)**



- At pressure < 5.2 atm:
- Phase change from solid → gas or vice versa
- When solid CO_2 heated at 1 atm, it sublimes at –78°C

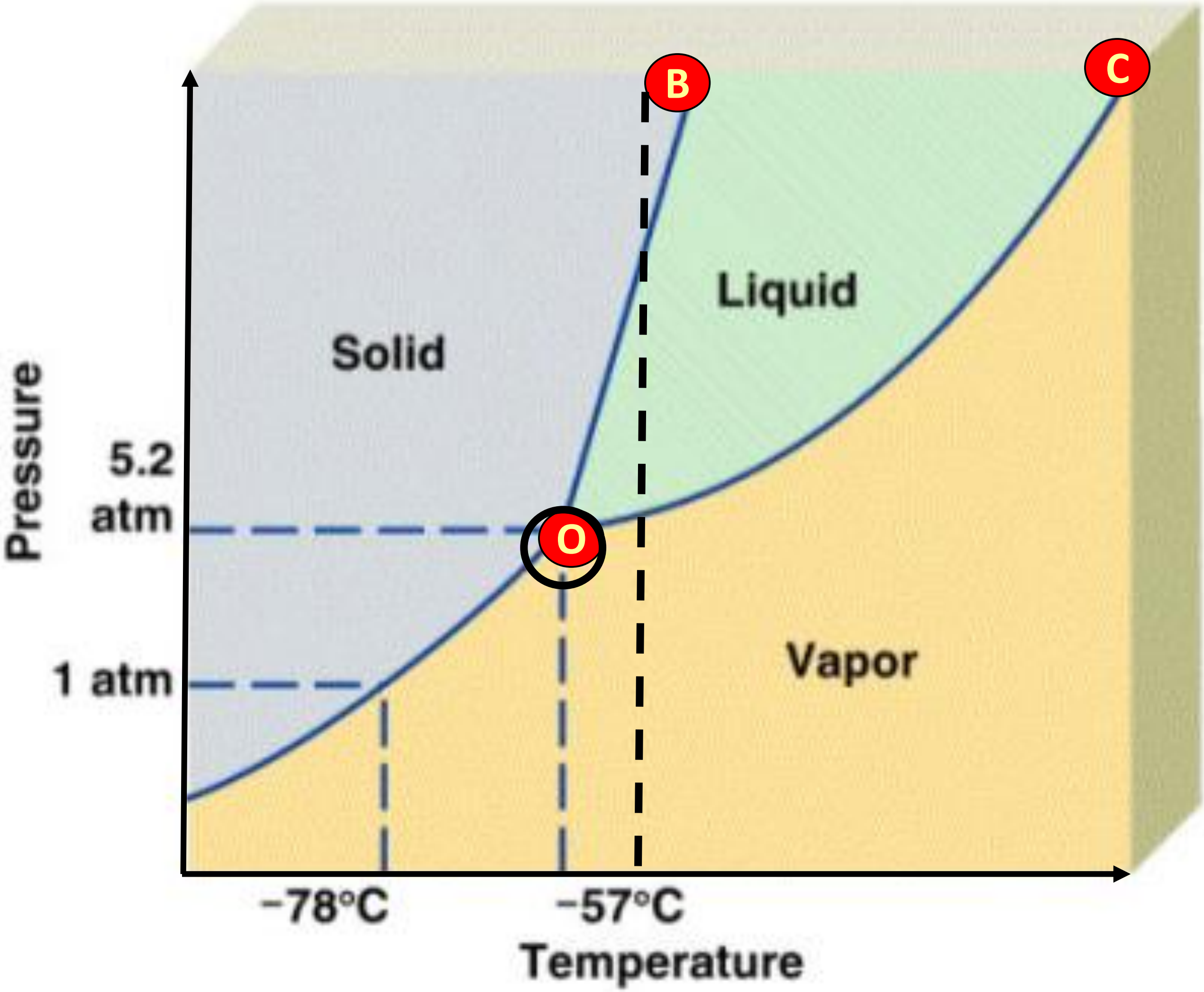


- At pressure > 5.2 atm:
- Phase change from solid → liquid → gas or vice versa

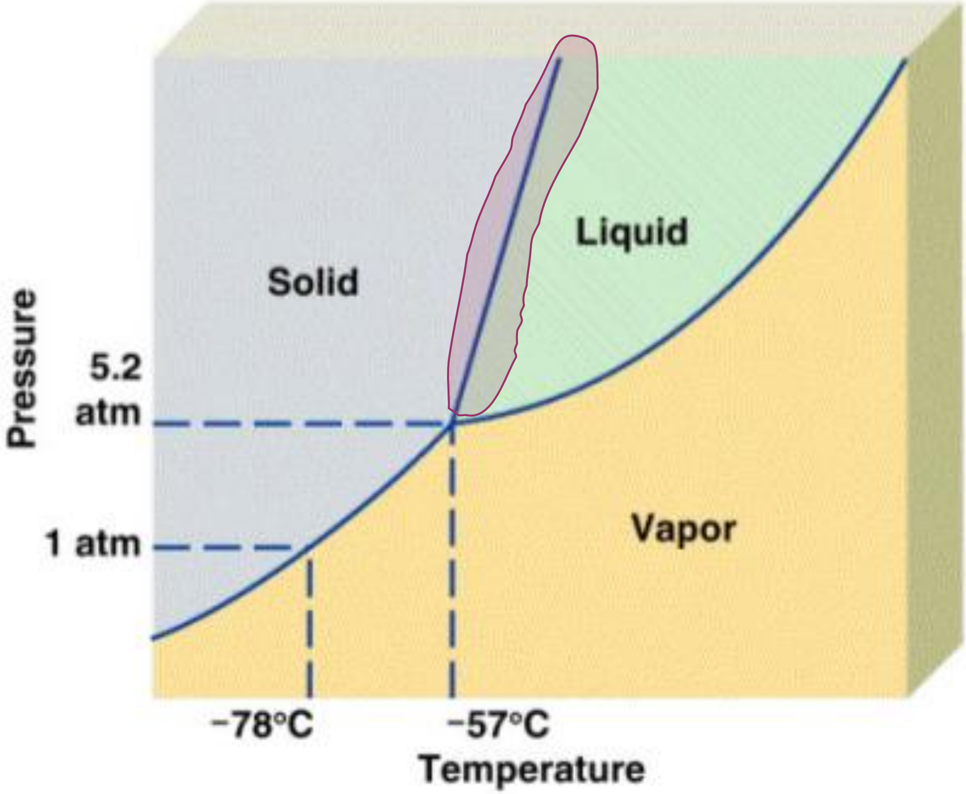


- At temperature < –57°C:
- Phase change from gas → solid or vice versa

(low *P*) (high *P*)



- At temperature **> –57°C**:
- **Phase change from gas → liquid → solid or vice versa**
  (low *P*) (high *P*)



The phase diagram for CO_2 is typical of most substances in that solid–liquid line slopes to the right (positive slope)

It means melting point (or freezing point) increase when the pressure increase

POSITIVE SLOPE

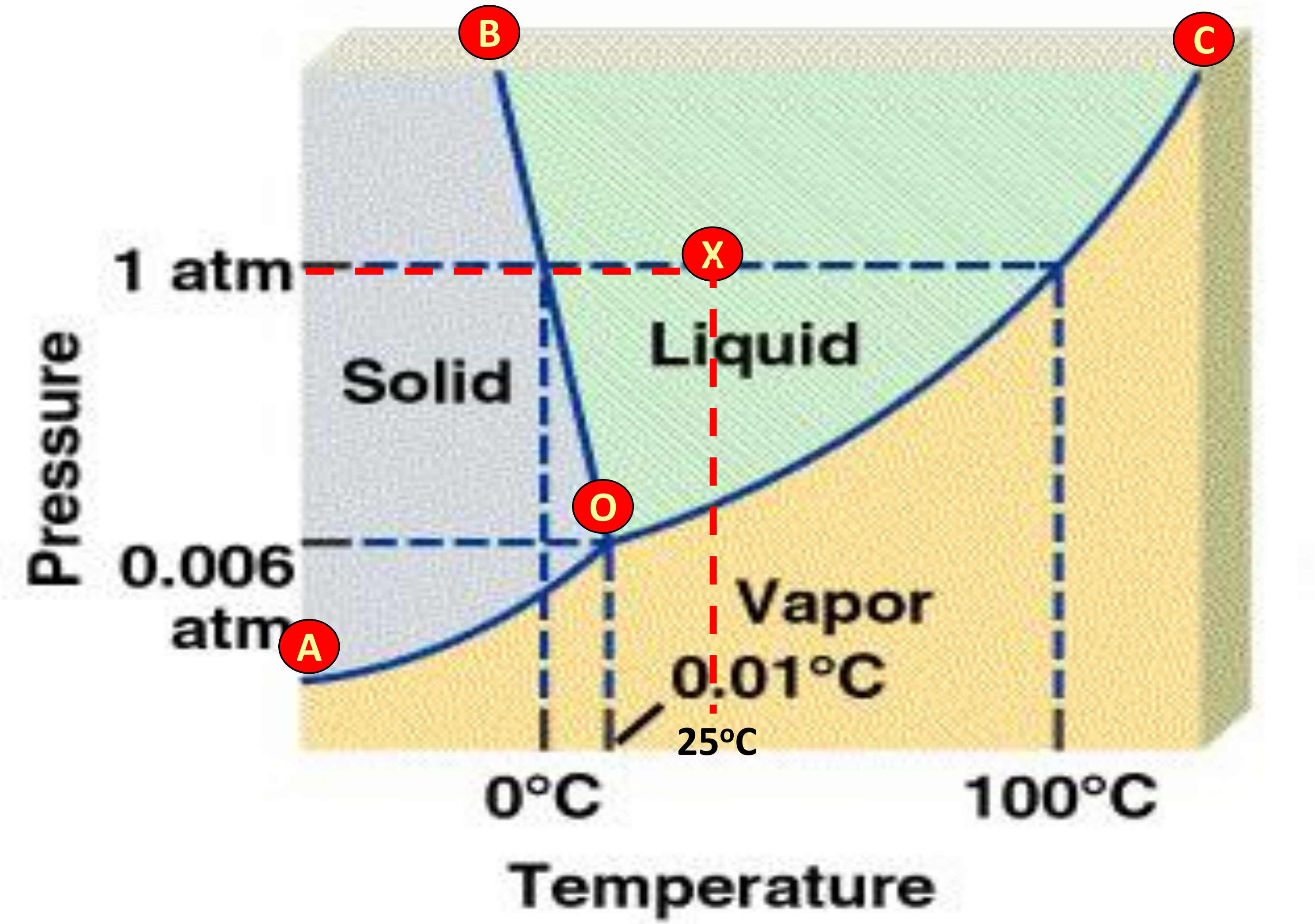
(Melting point of CO_2 increase when pressure increase)

Reason:

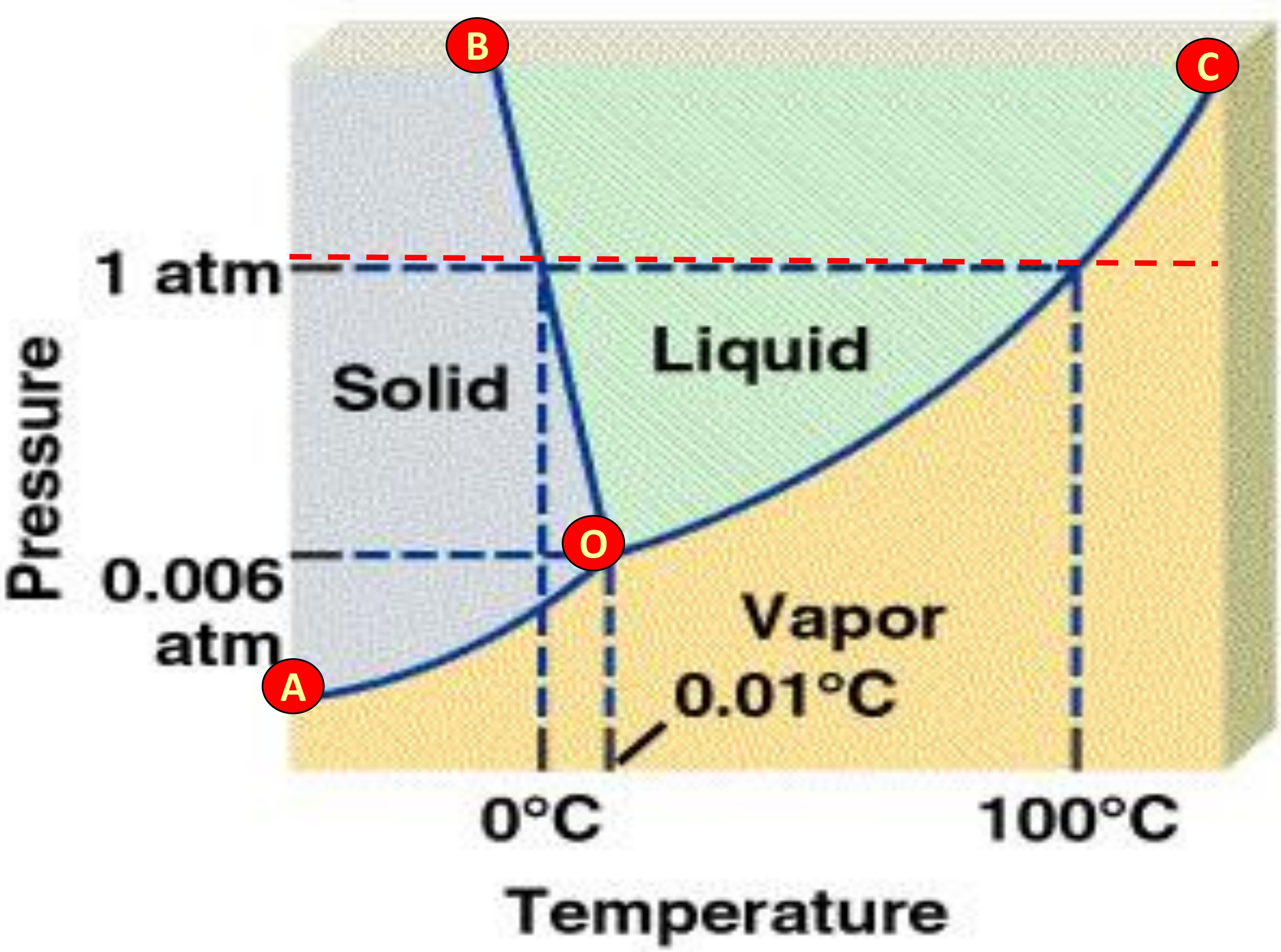
- Solid CO_2 is more dense (occupy smaller volume) than liquid CO_2.
- At high pressure, it favours the formation of solid CO_2 which has smaller volume.
- More heat is needed to melt the solid CO_2.

∴ solid CO_2 melts at higher temperature at high pressure.

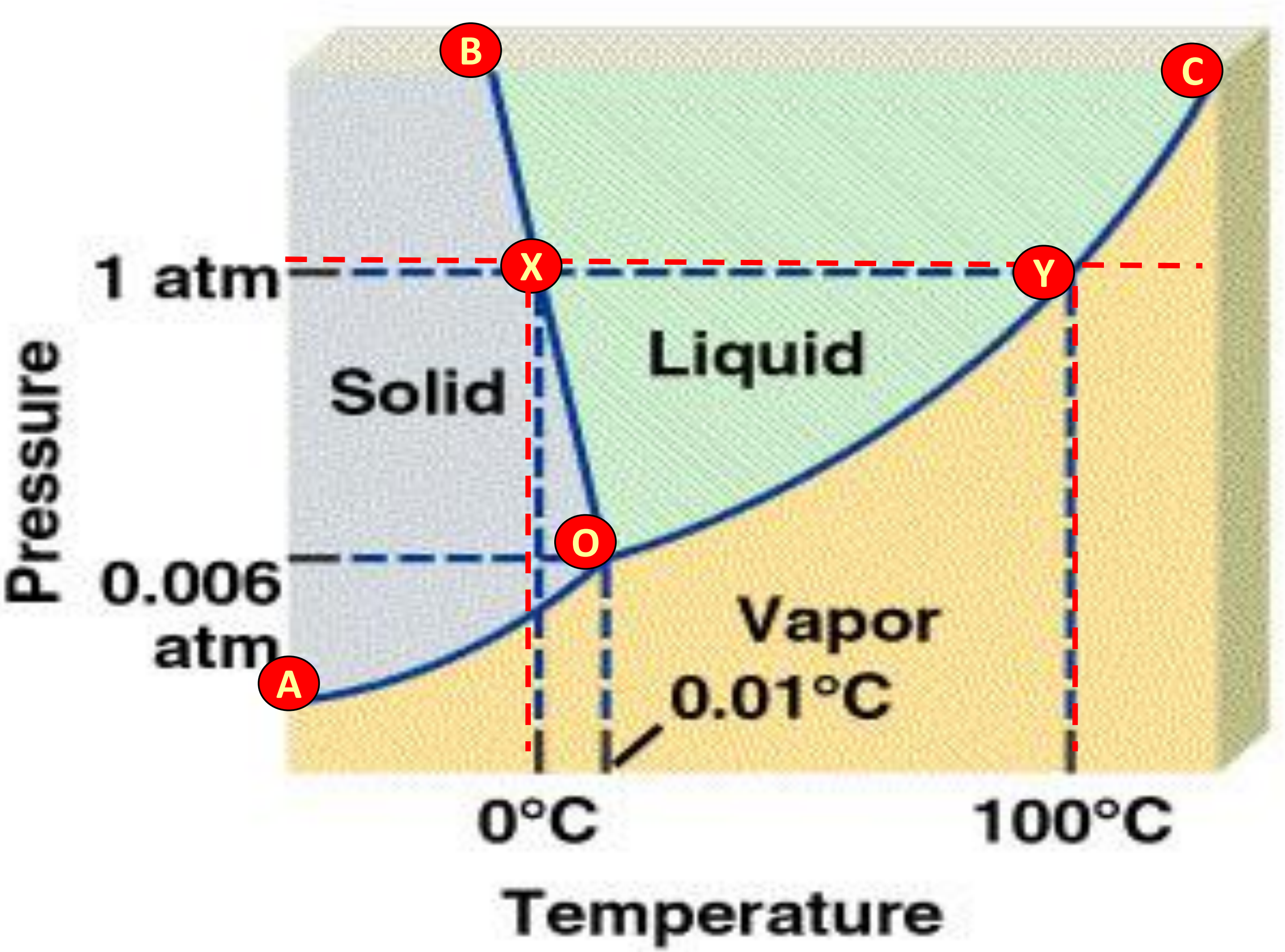
PHASE DIAGRAM OF H_2O



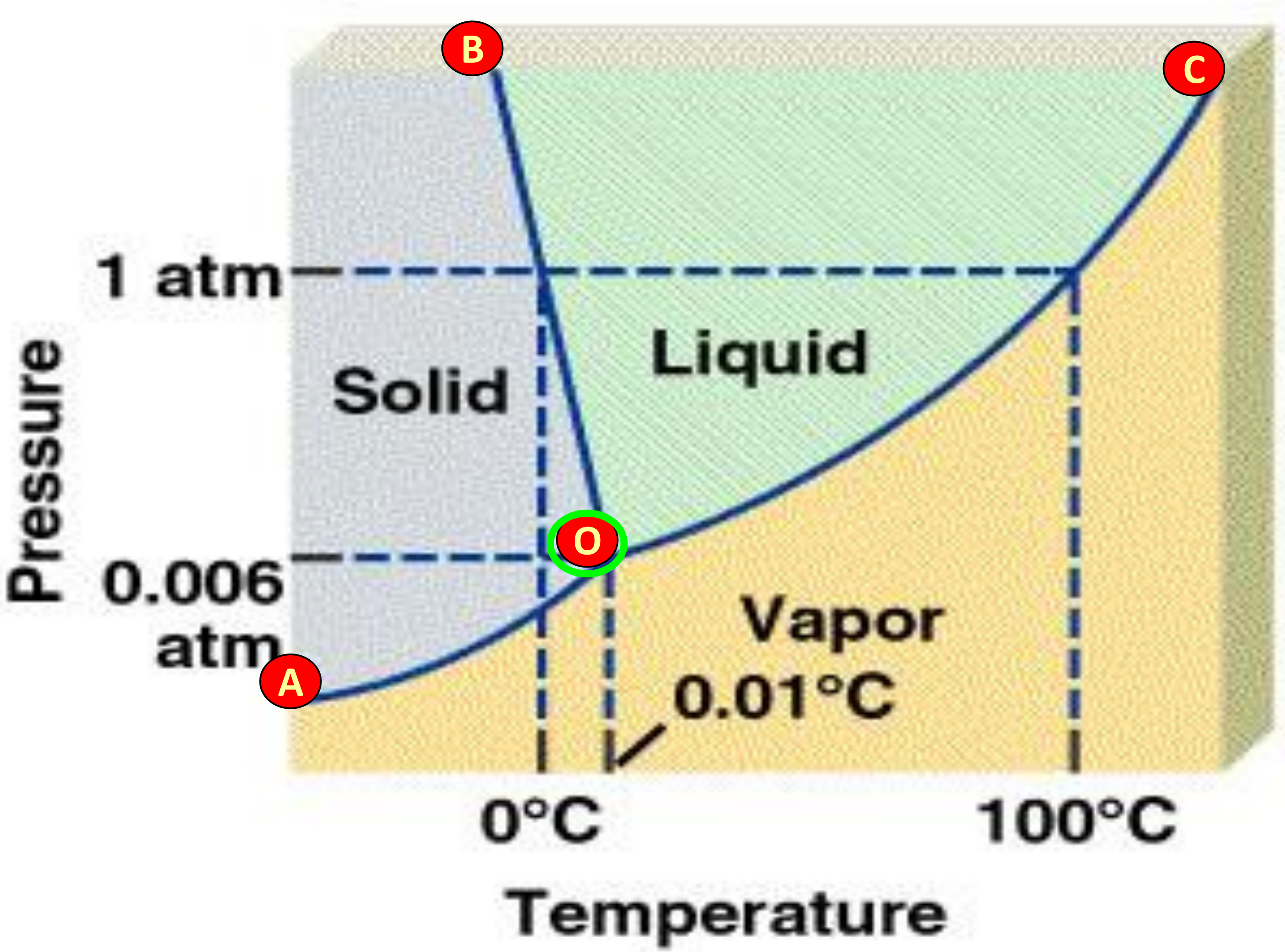
- H_2O is a liquid under normal conditions (25°C, 1 atm)



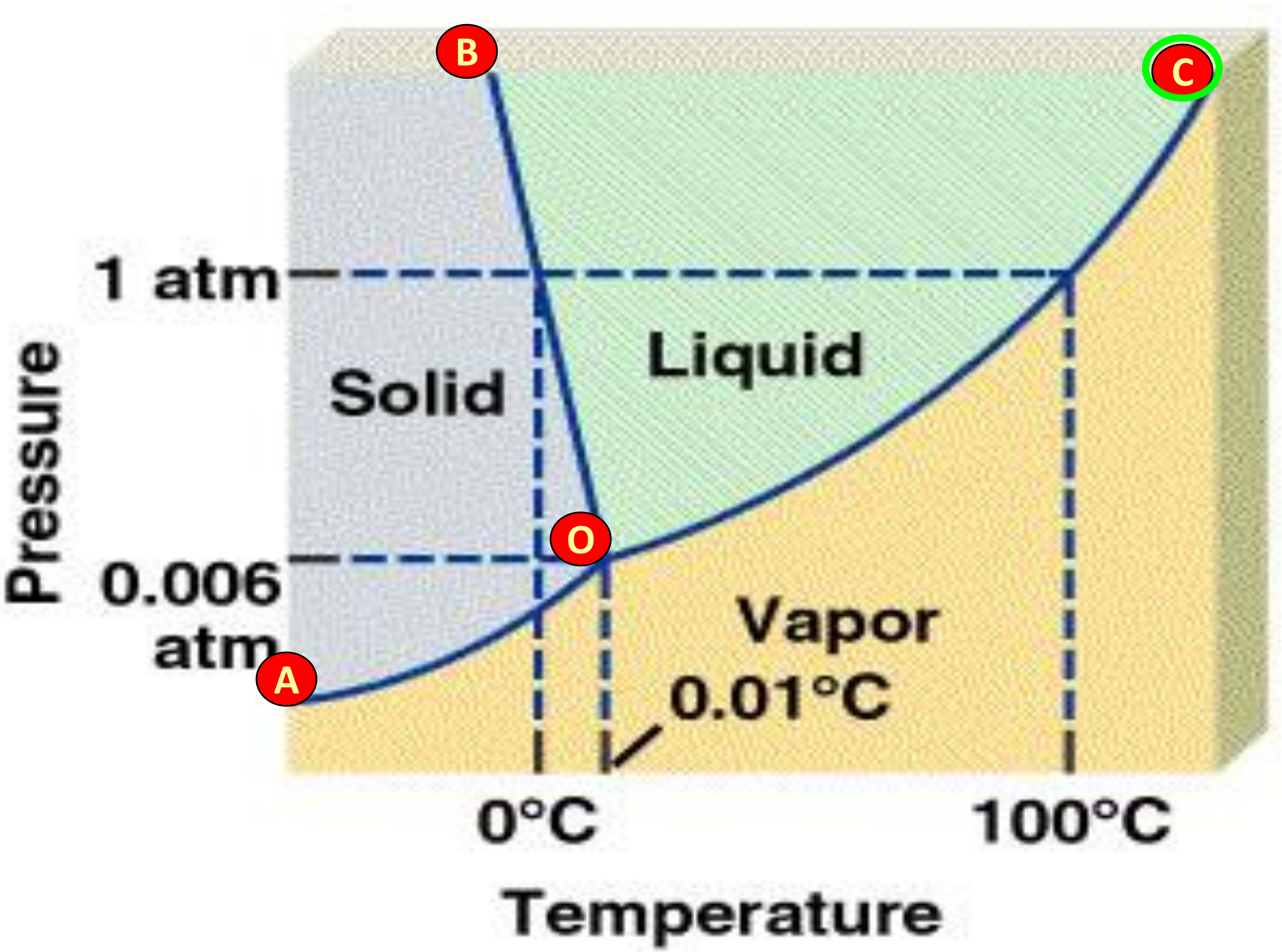
- At 1 atm, phase change from solid → liquid → gas or vice versa



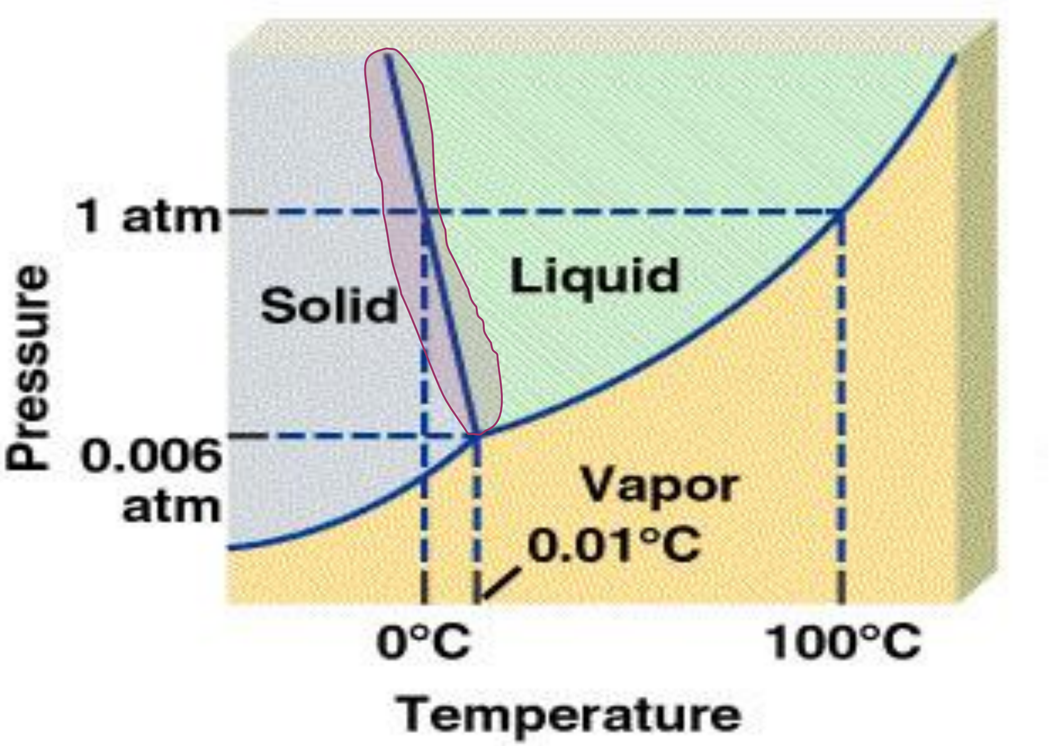
- At 1 atm, melting point / freezing point = 0°C boiling point = 100°C



- Point O = Triple Point (0.01ºC, 0.006 atm)



- **Point C = Critical Point (374ºC, 218 atm)**



The phase diagram for H_2O is one of few substances in that solid–liquid line slopes to the left (negative slope).

It means melting point decrease when pressure increase.

NEGATIVE SLOPE
(Anomalous behavior of H_2O)

- Melting point of decrease when pressure increase

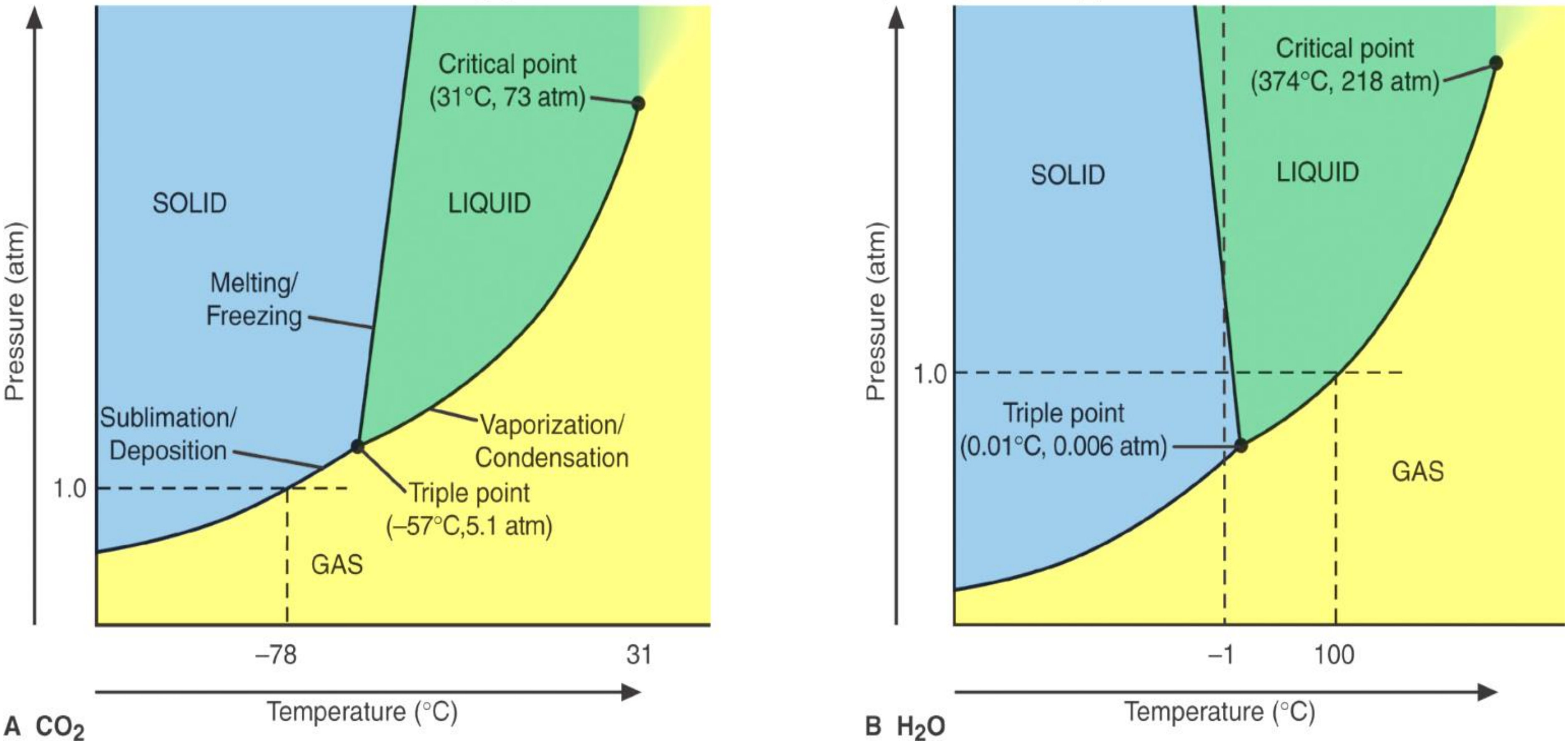
Reason:

- ➢ **Solid** H_2O ***(ice)*** **is less dense (occupy larger volume) than liquid** H_2O ***(water).***
- ➢ **At high pressure, it favours the formation of liquid** H_2O **which has smaller volume.**
- ➢ **Less heat is needed to melt the solid** H_2O.

solid H_2O **melts at lower temperature at high pressure.**

Phase diagram
CO_2 Vs H_2O

The phase diagram for **CO_2** has a solid–liquid curve with a positive slope, and the one for **H_2O** has a solid–liquid line with a negative slope. What macroscopic property can distinguish **CO_2** from **H_2O**?



- **If the solid is more dense than the liquid, the solid–liquid line slopes to the right; if less dense, to the left.**

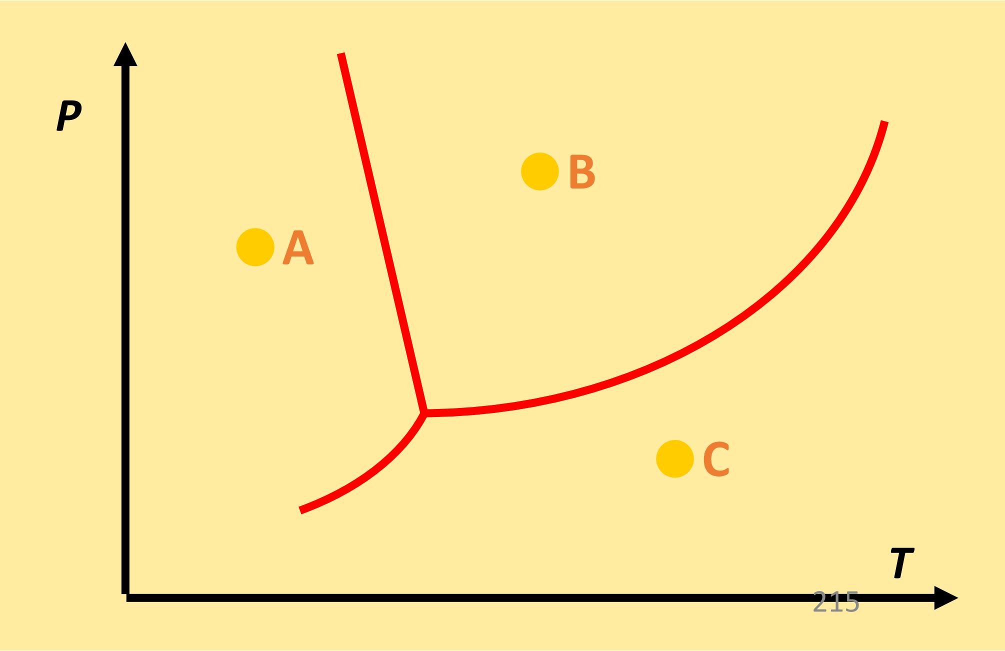
EXAMPLE 3 :

Predict what would happen as a result of the following changes: (State the phase changes)

(a) Starting at A, we raise the temperature at constant pressure.

(b) Stating at C, we lower, the temperature at constant pressure.

(c) Starting at B, we lower the pressure at constant temperature.

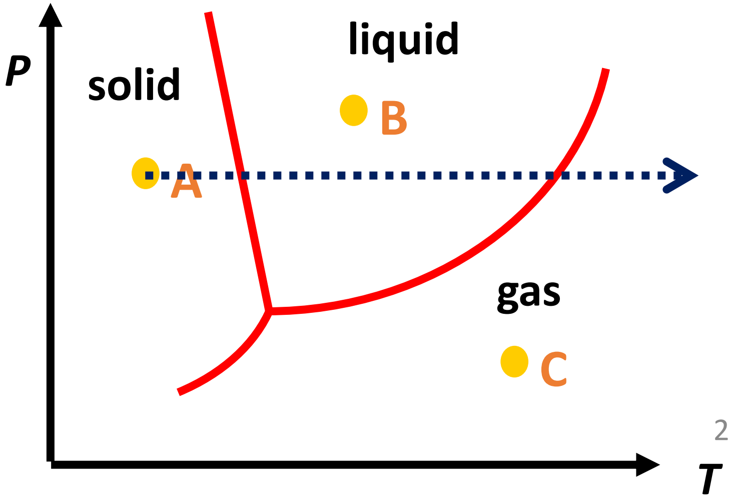


ANSWER 3:

(a) - Raising the temperature at constant pressure beginning at A implies starting with solid ice and warming until melting occurs.

- If the warming continued, the liquid water would eventually boil and change to steam.

- Further warming would increase the temperature of the steam.

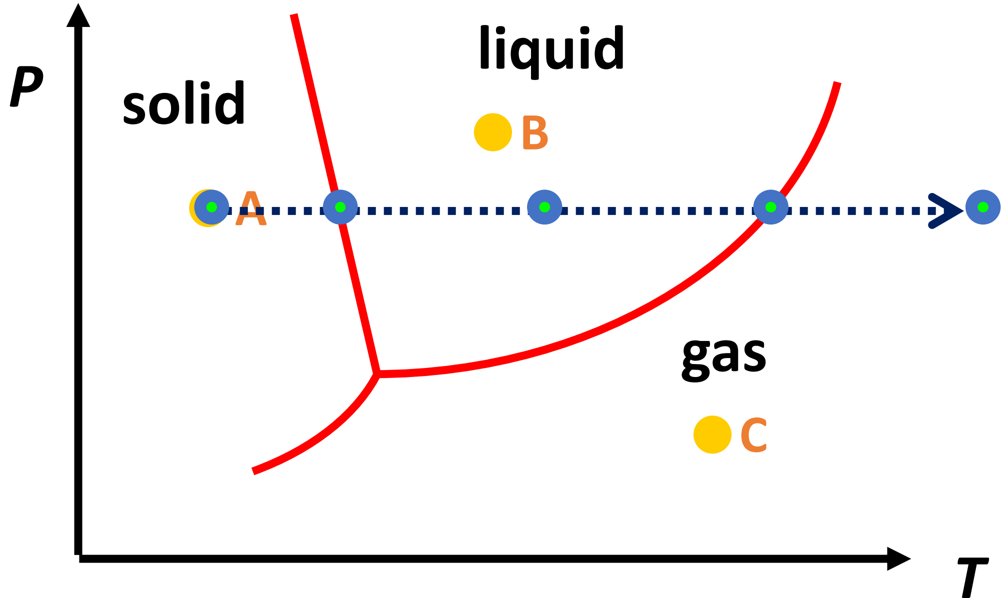


ANSWER 3:

(a)

Phase changes:

☞ **(solid) → (solid ⇌ liquid) → (liquid) → (liquid ⇌ gas) → gas**

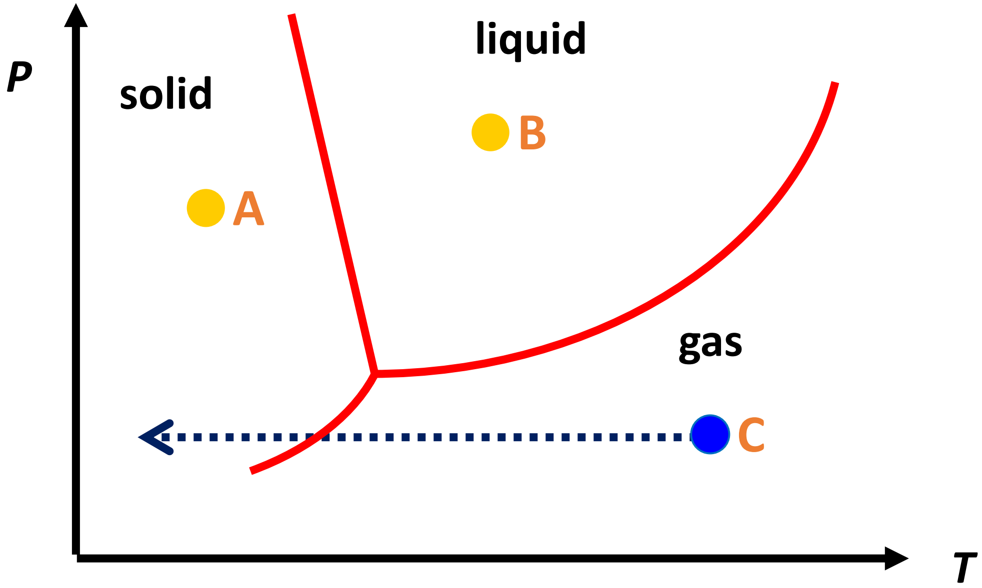


ANSWER 3:

(b) - At point C water is in the gas phase.

- Cooling without changing the pressure would eventually results in the formation of solid ice.

- Liquid water would never form.

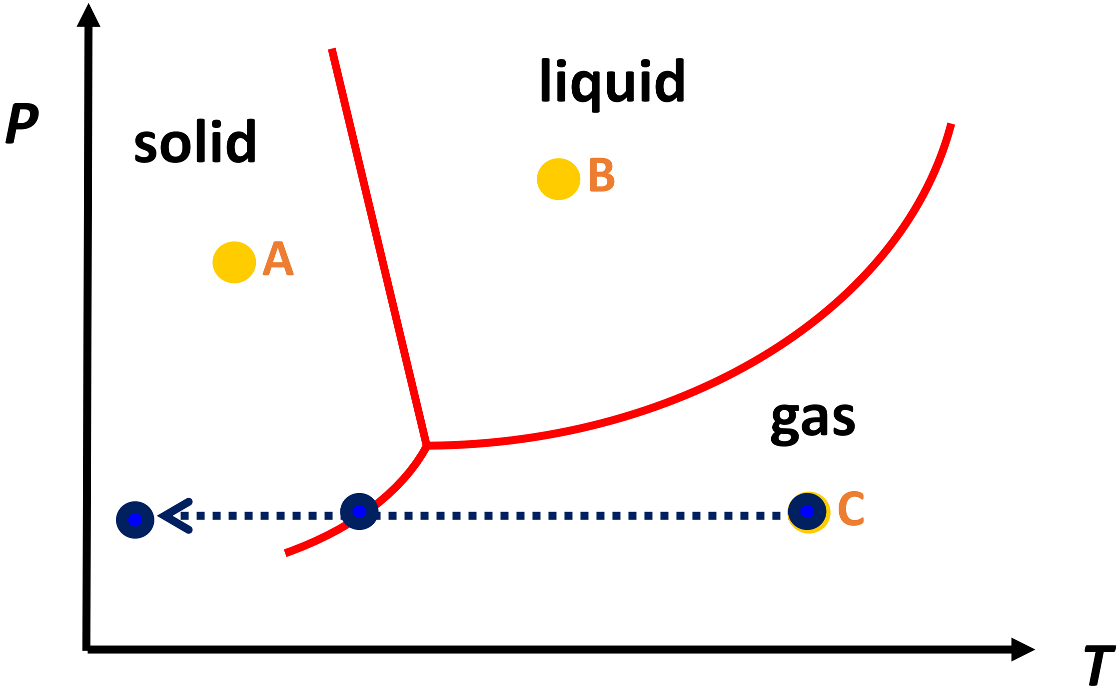


ANSWER 3:

(b)

Phase changes:

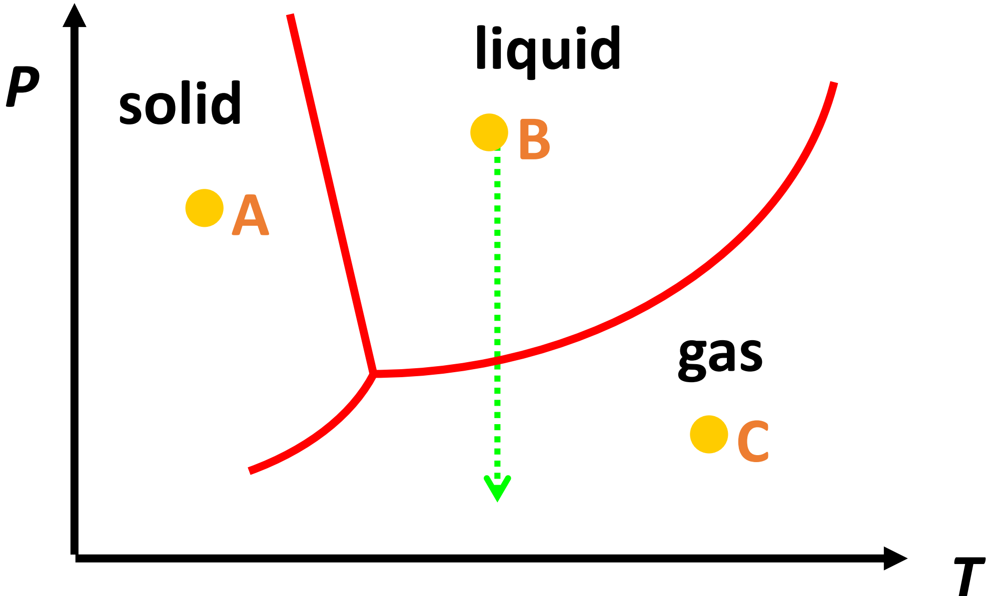
☞ (gas) → (gas ⇌ solid) → (solid)



ANSWER 3:

(c) - At B the water is the liquid phase.

- Lowering the pressure without changing the temperature would eventually result in boiling and conversion of water to the gas phase.

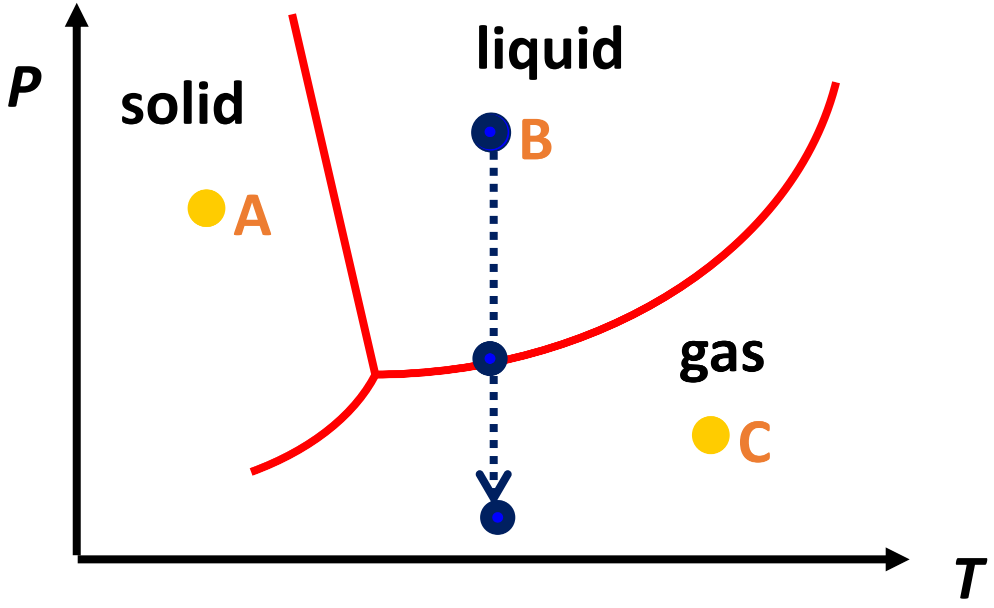


ANSWER 3:

(c)

Phase changes:

☞ (liquid) → (liquid ⇌ gas) → (gas)



Thanks!

For Attention

See You The Next Chapter

End Slide

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