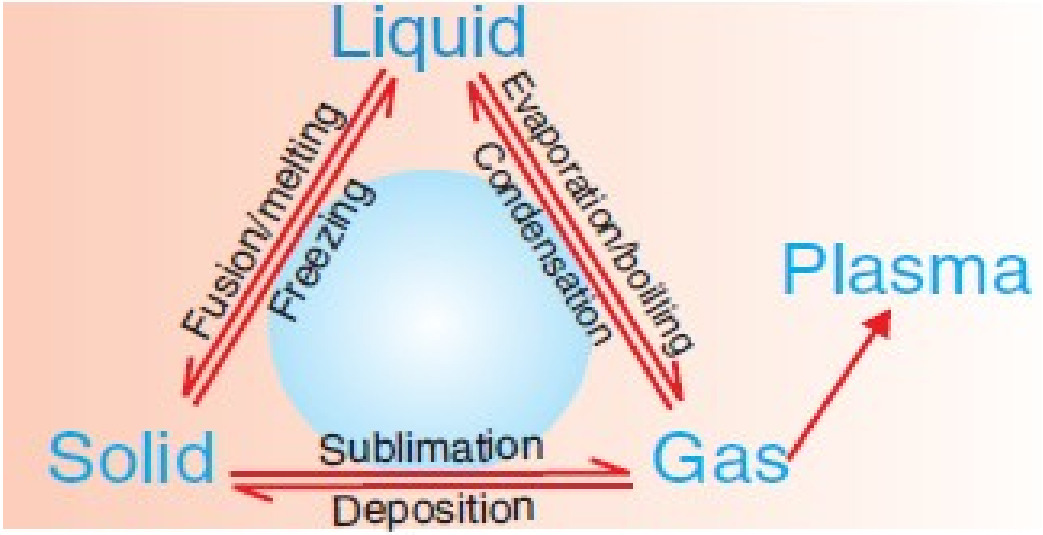
Course: Chemistry for Remedial Program

Chapter 3: Physical States of Matter



Unit Outcomes
After completing this unit,Students will be able to:

- ✓ understand the kinetic molecular theory and properties of the three physical states of matter;
- ✓ know the behavior of gases by using the variables volume, temperature, pressure and number of moles;
- ✓ know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor, pressure, boiling point, molar heat of vaporization,
- ✓ understand gas laws;
- ✓ develop skills in solving problems to which the gas laws apply;
- ✓ perform activities to illustrate gas laws;
- ✓ carry out experiments to determine the boiling points of liquids and the melting point of solids

3.1. Introduction

- You recall that matter is defined as anything that occupies space and has mass.
- In general, matter is known to exist in one of the three states called ***solid, liquid or gas.***
- ***Solids*** are orderly arranged arrays of atoms, molecules or ions having **definite volume** and **geometry** depending up on the arrangements of these atoms, molecules or ions and held together by strong attractive forces.
- **Liquids have definite volume**, but **do not have a definite shape**. Liquids take the shape of their container. In ***Liquid state***, molecules consisting of a liquid are held by molecular forces which are not as strong as solids.
- ***Gases*** have **neither a shape nor a volume** and the molecular forces of attraction are very much weaker.

❖ ***Fourth state is plasma state,*** a gaseous mixture of positive ions and electrons which occurs naturally in the interiors of stars with appreciable numbers of electrically charged particles.

- ✓Because of very high temperature (million degrees Celsius) and at such high temperatures molecules cannot exist, most or all of the atoms are stripped of their electrons.
- ✓ There fore, the extreme temperatures needed for fusion, no material can exist in the plasma state.

3.2.The kinetic theory of matter

- Kinetic theory of matter gives an explanation of the nature of the motion and the heat energy.
- According to the kinetic theory of matter, every substance consists of a very large number of very small particles called atoms, ions and molecules.
- The particles are in a state of continuous and random motion with all possible velocities.
- The motion of the particles increases with a rise in temperature.

Generaly, the kinetic theory of matter is based on the following assumptions:

- ✓All matter is composed of particles which are in **constant motion**.
- ✓The particles possess kinetic energy and potential energy.
- ✓The difference between the three states of matter is *due to their energy contents and the motion of the particles***.**

3.3. Properties of Matter

A. The properties of solids:

- They have mass
- They take up a place in space
- They have a definite volume and shape
- They cannot change their shape easily



There are intense attractive forces in solids that hold the molecules together, this is the reason why solids keep their shape and their volume and they can only vibrate

Cont..

- In solids the *cohesive forces* are much greater than the *disruptive forces.*
- ✓In a crystalline solid, each molecule is held at a fixed location within a

crystal lattice.

- ✓The disruptive forces cause the molecules to vibrate at their fixed location, but *does not allow the molecules to move past one another*.

The kinetic-molecular model of matter can explain the characteristic properties displayed by solids.

1. High density

Because the cohesive forces are dominant, the molecules in a solid are very close together, which results in a greater number of molecules in a given volume.

2. Definite shape

Because the molecules in a solid *cannot move past one another*, they maintain a defined shape.

3. Low compressibility

Because the molecules in a solid are very close together, increasing the pressure is unable to move them much closer together.

4. Small thermal expansion

Though heating causes the molecules in a solid to vibrate faster about their fixed positions, the cohesive force still predominates, holding them very close together

Classes of Solids based on their structural features
(on the arrangement of constituent particles)

i. Crystalline Solids

- What are Crystalline Solids?

The solids featuring highly ordered arrangements of their particles (atoms, ions, and molecules) in microscopic structures are *called crystalline solids.*

Types of Crystalline Solid

Crystalline solids can be further classified depending on the nature of the bonding as:

1. *Ionic (Nature of force: Strong electrostatic forces of attraction) Eg: NaCl, LiF, MgO, etc.*
2. *Molecular, contain discrete molecular units held by relatively weak intermolecular forces (VanderWaal'sforces, Dipoleinteraction, Hydrogen bonding). Eg: Solid Ar, Kr and etc.*
3. *Covalent (Covalent bonds). Lattice points are atoms. Eg: Diamond,*

ii. Amorphous Solids

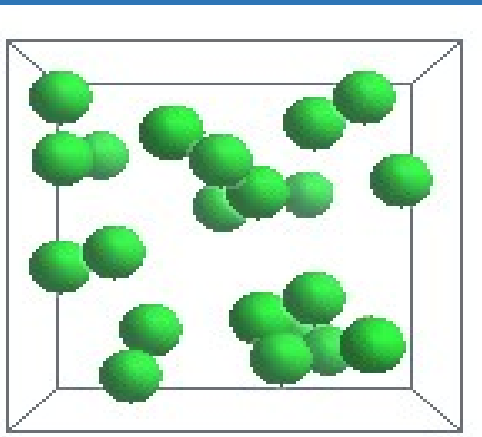
- What are Amorphous Solids?
- The solids in which the particles are ***not arranged in any specific order*** *or the solids* that **lack the overall order of a crystal lattice** are called amorphous solids.
- The term '**amorphous',** when broken down into its Greek roots, can be roughly translated to **"*without form*".** Many polymers are amorphous solids.
- Examples of such solids include glass, gels, and nanostructured materials.

- **An ideal crystal** is defined as an atomic arrangement that has infinite translational symmetry in all the three dimensions, whereas such a definite definition is not possible for an ideal amorphous solid (a-solid).

B. The properties of liquids:

- They have mass
- They take up a place in space
- They have a definite volume But, does not have shape
- They have higher densities than gases.
- They take the shape of their container
- They are fluids and A fluid is a substance that can easily





In the liquid state the cohesive forces still predominate, holding the molecules almost as close together as in the solid state.

- ✓ *The disruptive forces are strong enough, though, that the molecules do not have fixed locations, they instead are able to move about, slipping past one*

The kinetic-molecular model of matter can explain the characteristic properties of liquids.

1. *Because the cohesive forces are able to still hold the molecules close together in liquid, their densities, compressibility and thermal expansions*

   are similar to their corresponding solids, because all three of these properties depend on the distances separating the molecules.

2. *Unlike solids, liquids do not have defined shapes because, unlike solids their molecules are capable of moving past one another.*

Superficial tension and Viscosity

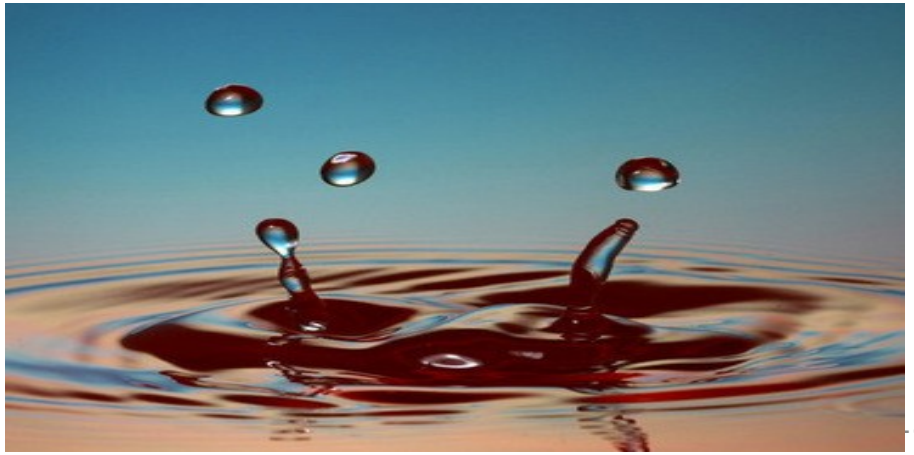
- **There are attractive forces in liquids (but less intense than in solids) that hold the liquid molecules together.**
- **These attractive forces prevent that the molecules from separating but not from moving relative to each other.**
- **This is the reason why liquids keep their volume but do not hold their shape.**

Viscosity

- ✓ Is the resistance of a liquid to flow.
- ✓ Oil is less dense than water because it floats on water but it is more viscous because the oil slides worse on a surface.

Superficial tension

- ✓ Is the characteristics of liquid that is the resistance to the penetration of bodies in it.
- ✓ One of the substances that has greater superficial tension is water.
- ✓ For that reason, it is possible that some insects walk on water. The superficial tension is the cause of the spherical form of the drops of liquids.
- ✓ Another consequence of the superficial tension is the ascent of liquids within tubes of small diameter.



Surface tension of a liquid

- Is defined as the *force per unit length* acting perpendicular on an imaginary line drawn on the liquid surface, tending to pull the surface apart along the line.

$$T = \frac{F}{L}.$$

- **where, T -Surface tension, F- force, L- length**
- Its unit is N m^{-1}
- It depends on temperature.
- The surface tension of all liquids decreases linearly with temperature

Viscosity

- **Viscosity is the property of the fluid** by virtue of **which it opposes** relative motion between its different layers.
- Both liquids and gases exhibit viscosity but **liquids are much more viscous than gases.**

Example

- *Water, milk, and fruit juice all flow very easily*; this can be observed when you pour each into a glass.
- These are all examples of ***low viscosity or thin liquids***.
- Other examples such as vegetable oil, maple syrup and dish soap are considerably **more viscous,** resisting flow and pouring out more slowly

- **If F is the viscous force on the layer then , F ∝ A, where A is the area of the layer and**

$$F = -\Delta v / \Delta t$$

- The negative sign is put to account for the fact that the viscous force is opposite to the direction of motion Thus;

$$F = -\eta A \frac{dv}{dt}$$

Where η is a constant depending upon the nature of the liquid and is called the coefficient of viscosity and

$$velocity\ gradiant = \frac{dv}{dt}$$

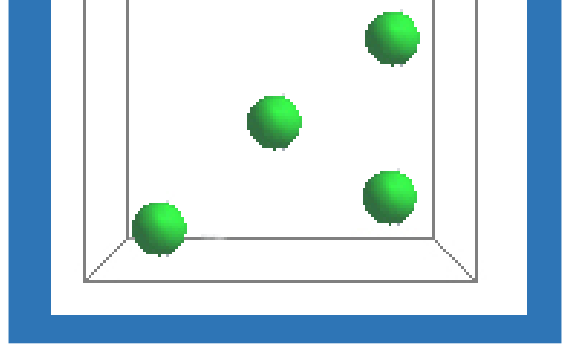
If A = 1 and dv/dx = 1. We have F=- η

Thus the coefficient of viscosity of a liquid may be defined as the viscous force per unit area of the layer where velocity gradient is unity

C. The properties of gases:

- They have mass
- They take up a place in space
- They take up all of the ava space
- They have *No shape and Volume*
- They mix readily with other gases
- They can compress into a small space and they can expand

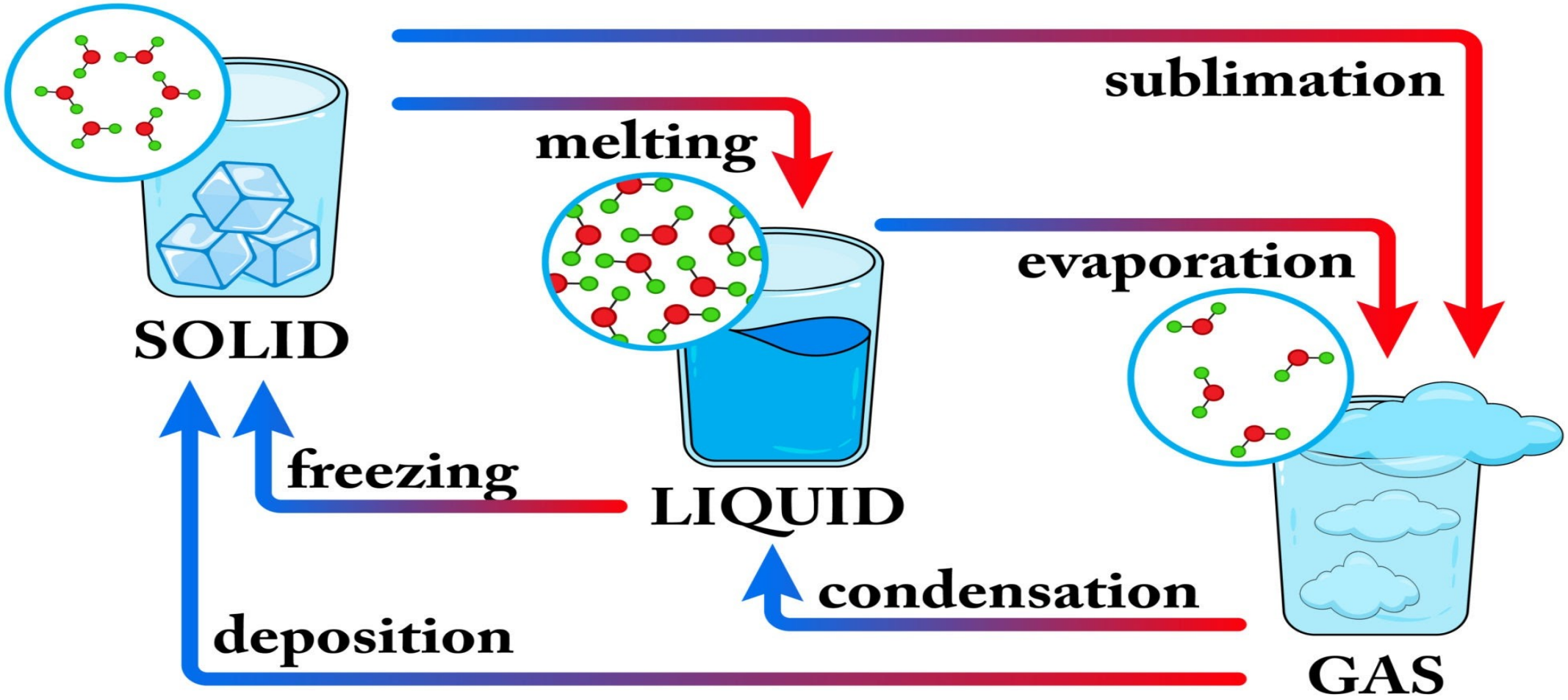
- **There are not attractive forces in gases.**
- **This is the reason why molecules of gases are very distant, can move freely in any direction and they don't keep their shape, and don't keep their volume either.**



kinetic-molecular model of matter can explain the characteristic properties displayed by gases.

1. ***Low density –*** because the disruptive forces are dominant, the molecules are spread out and as far apart from one another as possible. This give gases very low densities, and therefore little mass in a giving volume
2. **Indefinite shape –** Like liquids, the molecules can easily move past one another. They take on the shape of their container.
3. ***Large compressibility –*** In gases the molecules are spread out with lots of space between them. When pressure is applied they can easily be pressed closer together.
4. ***Moderate thermal expansion*** – as a gas is heated the individual molecules acquire more velocity, which means they strike the walls of the container with a greater force

D. Changing states of matter



CHANGE OF STATE	WHAT IS IT?
Melting	It is the change from a solid to a liquid.
Freezing	It is the change from a liquid to a solid.
Vaporization	It is the change from a liquid to a gas, to the temperature of boiling and in all the mass of the liquid.
Evaporation	It is the change from a liquid to a gas, to any temperature and in the surface of the liquid.
Condensation	It is the change from a gas to a liquid.
Sublimation	It is the change from a solid to a gas.
Deposition	It is the change from a gas to a solid.

3.4.1. The kinetic molecular theory of gases Assumptions;

i. Gas particles are in a state of constant, continuous, rapid, random motion and, therefore, **possess kinetic energy**.

- The motion is constantly interrupted by collisions with molecules or with the container.
- The pressure of a gas is the effect of these molecular impacts.

ii. The volume of the particles is negligible compared to the total volume of the gas.

- Gases are composed of separate, *tiny invisible particles called molecules.*
- Since these molecules are so far apart, the total volume of the molecules is extremely small compared with the total volume of the gas.
- Therefore, under ordinary conditions, gas consists of empty space.
- This assumption explains why gases are so easily compressed and why they can mix so readily.

iii. *The attractive forces between the particles are negligible*.

There are no forces of attraction or repulsion between gas particles.

You can think of an ideal gas molecule as behaving like small billiard balls.

When they collide, they don't stick together but immediately bounce apart.

iv. *The average kinetic energy of gas particles depends on temperature of gases.*

At any particular moment, the molecules in a gas have different velocities.

The mathematical formula for kinetic energy is:

$$K.E. = \frac{1}{2} mv^2$$

where *m* is mass and *v* is velocity of gas molecules.

- Because the molecules have different velocities, they have different kinetic energies.
- However, it is assumed that the average kinetic energy of the molecules is directly proportional to the absolute *(Kelvin)* temperature of the gas.

3.4.2. The Gas Laws

- The gas laws are the products of many experiments on the physical properties of gases, which were carried out over hundreds of years ago.
- Simple mathematical equations can be derived that relate a gas's volume, pressure and temperature.

$$PV = nRT$$

Ideal Gas Equation

- These equations are *called state equations* because they describe mathematical relationships between the ***volume, temperature, pressure, and quantity of a gas (number of moles.)***

a. Pressure:

Pressure is defined as the force applied per unit area.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

The SI unit of pr [illegible] :ton per square meter. $1\ \text{Pa} = 1\ \text{N/m}^2$ and

$$1\ \text{atm} = 760\ \text{mmHg} = 76\ \text{cmHg} = 760\ \text{torr} = 101325\ \text{Pa} = 101.325\ \text{kPa}$$

b. Volume:

Volume is the space taken up by a body. The SI unit of volume is the cubic metre (m^3). Other common units of volume are millilitre **(mL)** and litre **(L).**

A litre is equivalent to one cubic decimeter: The relation is given as follows;

1 L = 1000 mL = 1000 cm3 =

$$1\ \text{cm}^3 = (1\times10^{-2}\ \text{m})^3 = 1\times10^{-6}\ \text{m}^3$$

$$1\ \text{dm}^3 = (1\times10^{-1}\ \text{m})^3 = 1\times10^{-3}\ \text{m}^3 = 1\ \text{L}$$

C. Temperature:

- ✓ Temperature is the degree of hotness or coldness of a body.
- ✓ Three temperature scales are commonly used. These are °F *(degree Fahrenheit), °C (degree Celsius) and K (Kelvin)*.
- ✓ In all gas calculations, we use the ***Kelvin*** scale of temperature.
- ✓ We use the following formulae for conversions:

$$^{\circ}C = (^{\circ}F - 32)\frac{5}{9}$$

K = °C + 273

$$^{\circ}F = \left(\frac{9}{5} \times {}^{\circ}C\right) + 32$$

Molar Volume and Standard Conditions (STP)

- The conditions of a pressure of 1 atmosphere and a temperature of **0°C** *(273.14 K)* are called standard temperature and pressure or STP for gases.
- At STP the volume of one mole of any gas is equal to *22.4* litres.
- This volume is *known as molar gas volume*.

d. Quantity of gas:

- ✓ The quantity of a gas is expressed in *mole (n).*
- ✓ **Mole:** is the quantity of gas in terms of number of particles.
- ✓ It is the number of atoms or molecules in 1 gram-atom or 1 gram-molecule of an element or a compound.

The gas laws developed from:

1. Boyle's Law

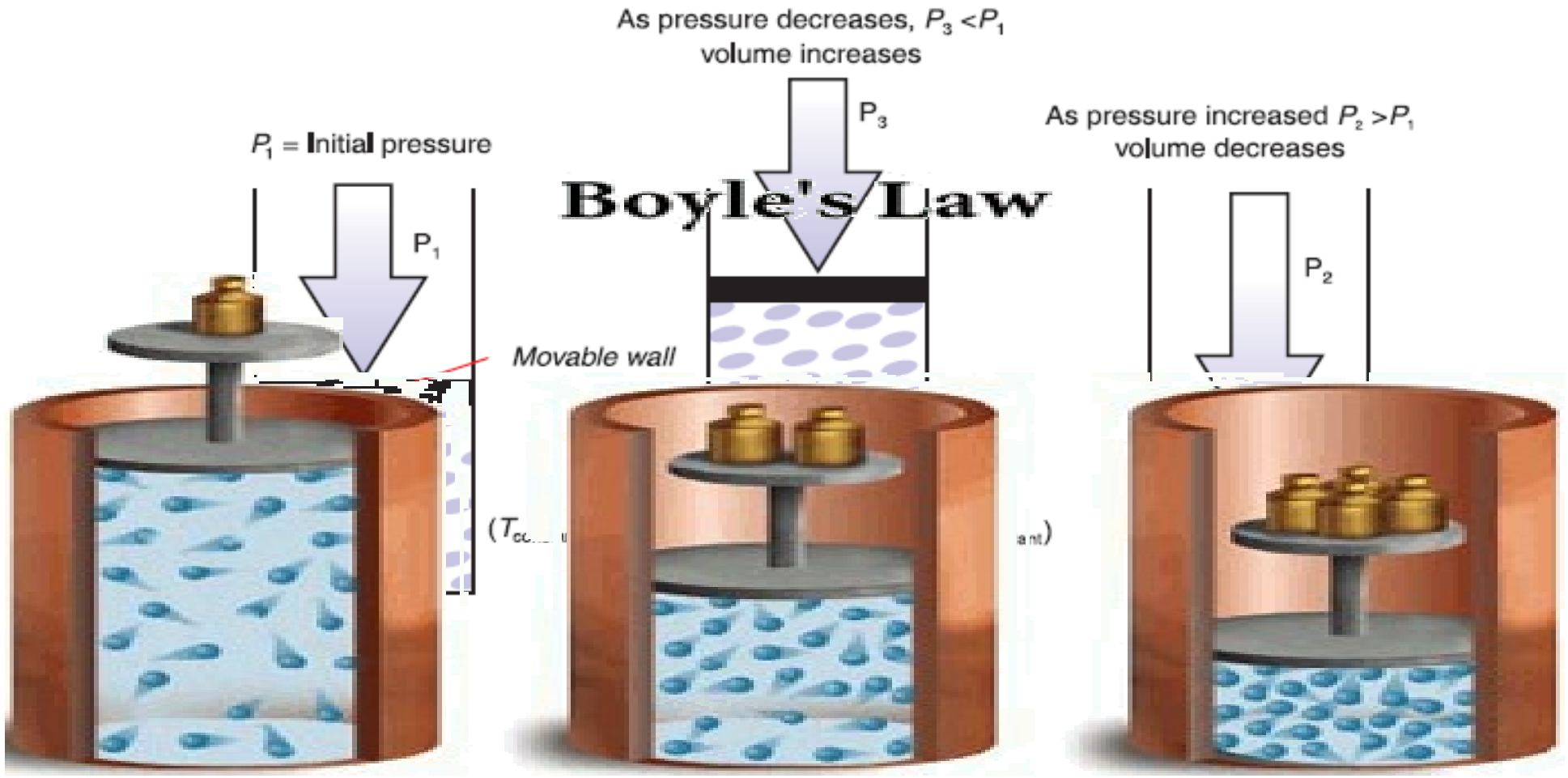
Robert Boyle (1627-1691):

- His experiment helped to analyze the relationship between the volume and pressure of a fixed amount of a gas at constant temperature.
- Decreasing the external pressure, causes the gas to expand and to increase in volume.
- Correspondingly, increasing the external pressure allows the gas to contract and decrease in volume.

***Boyle* studied** the relationship between the pressure of the trapped gas and its volume.

- ✓ Accordingly, he discovered that at constant temperature doubling the pressure on a sample of gas reduces its volume by one-half.
- ✓ Tripling the gas pressure reduces its volume to one-third of the original.

Figure: The relation between pressure and volume



- **Boyle's law states that** *the volume of a fixed mass of gas is inversely proportional to the pressure at a constant temperature.*
- *Mathematically given as:*

$$V \ \alpha \ \frac{1}{P} \ (\text{at constant } T \text{ and } n)$$

From which follows,

$$V = k\frac{1}{P} \quad \text{or} \quad PV = k;$$

where k is a constant at a specific temperature for a given sample of gas.

- *If* $\boldsymbol{P_1}$ ***and*** $\boldsymbol{V_1}$ *represent the initial conditions; and* $\boldsymbol{P_2}$ ***and*** $\boldsymbol{V_2}$ *represent the new or final conditions, Boyle's law can be written as:*

$$\boldsymbol{P_1V_1 = P_2V_2}$$

Example 1

An inflated balloon has a volume of 0.55 L at sea level (1.0 atm) and is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm.

- Assuming that the temperature remains constant, what is the *final volume of the balloon?*

Solution:

Givens: *Initial conditions* *Final conditions*

$P_1 = 1.0$ atm $P_2 = 0.40$ atm

$V_1 = 0.55$ L $V_2 = ?$ $V_2 = P_1V_1/P_2$

$V_2 = 0.55L \times 1/0.4 = 1.4L$

Example 2

A certain gas occupies a volume of 10.0 m^3 at a pressure of 100.0 kPa. If its volume is increased to 20 m^3, what would be the new pressure of the gas assuming temperature remains constant?

Solution:

Given: $V_1 = 10.0 m^3$ $V_2 = 20.0 m^3$

$P_1 = 100.0$ kPa $P_2 = ?$

$$P_2 = P_1V_1/V_2$$

$$P_2 = 1/2 \times 100.0 \text{ kPa} = 50 \text{kPa}$$

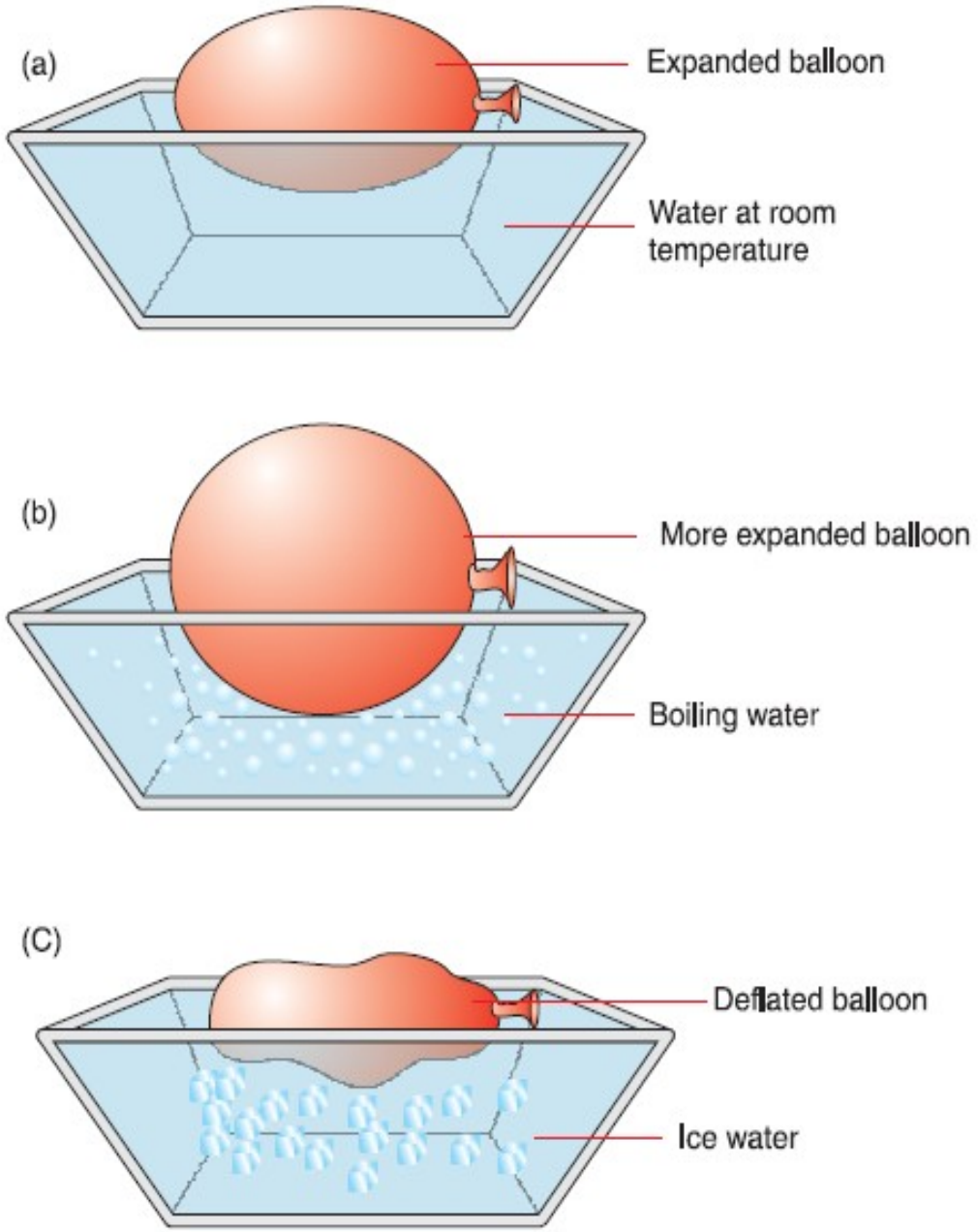
2. Charles' Law

- The French physicist, **Jacques Charles** (1746-1823), was the first person to fill **a balloon with hydrogen gas** and made the first solo balloon flight.
- **Charles** investigated *quantitative relationship between the volume and temperature* of a fixed quantity of gas which is held at constant pressure.
- **He** stated that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.
- In 1848, *Lord Kelvin* realized that a temperature of -273.15°C is considered as **absolute zero. It is theoretically the lowest attainable temperature.**
- Then he set up an absolute temperature scale, or the Kelvin temperature scale, with **absolute zero as the starting point on the Kelvin scale**.

For example,

- Doubling the Kelvin temperature causes the volume of a gas to double, and
- Reducing the Kelvin temperature by half causes the volume of a gas to decrease by half.

- This relationship **between Kelvin temperature and the volume** of a gas is ***known as Charles' law.***



Mathematically:

$V \propto T$ at constant P and n:

$$V = kT, \quad \text{or} \quad \frac{V}{T} = k$$

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

Figure: Relationship between the volume of air in the balloon and its temperature

3. Combined Gas Law

Combined Gas Law

1. Boyle and Charles' Laws can be combined to give the combined gas law equation:

$$\frac{PV}{T} = k''$$

2. Initial and final states
   a. If a gas is changed from and initial (*i*) to a final (*f*) state, then there will be a set values for P, V, and T, for each state:
      i. Initial state -. P_I, V_I, T_i
      ii. Final state - P_f V_f, T_f
   b. Applying the combined gas law to each:
      i. Initial state: $\frac{P_i V_i}{T_i} = k''$
      ii. Final state: $\frac{P_f V_f}{T_f} = k''$
   c. Since both are equal to *k''*, we can set them equal to each other:

$$\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$$

Combination of Boyle's, Charles's and Avogadro's Laws give the Ideal gas law equation.

$$\frac{PV}{T} = k''$$

$$\frac{PV}{T} = nk'''$$

$$\frac{PV}{T} = nR \quad (Ideal\ Gas\ Law\ Equation)$$

$$PV = nRT$$

1. Where $R = k''''$ is the *Ideal Gas Law Constant* or *Universal Gas Constant*
   a. $R = 0.0821\ L \bullet atm/mol \bullet K$
   b. RT give the kinetic energy per mole of gas
2. An ideal gas is one that has no potential energy
   a. It is neither attracted to the itself or to the walls of the container

Exercise:

A 200 mL sample of oxygen gas is collected at 26°C and a pressure of 690 Torr. What volume will the gas occupy at STP (0°C and 760 Torr)?

4. Dalton's Law

Dalton's Law

A. For ideal gases (no potential energy) at a fixed volume and temperature, the pressure exerted by the gas on the walls of the container is proportional to number of moles of gas present:

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

B. Dalton's Law of partial pressures states that for a mixture of gases at a fixed volume and temperature, each gas will contribute to the total pressure in proportion to its number:

$$P_{tot} = \frac{n_A RT}{V} + \frac{n_B RT}{V} + \frac{n_C RT}{V} + \ldots$$

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

1. The contribution made by each gas is called its partial pressure
2. In lab, you made use of this relationship in the gas laws experiment when you subtracted the partial pressure of water from the total pressure to get the partial pressure of the CO_2.

5. Graham's Law

- Graham's law of diffusion states that *at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.*

Mathematically it can be expressed as:

$$r \propto \sqrt{\frac{1}{d}} \quad \text{or} \quad r \propto \sqrt{\frac{1}{M}}$$

For two gases (Gas 1 *and* Gas 2), their rates of diffusion can be given as:

$$r_1 \propto \sqrt{\frac{1}{d_1}} \quad or \quad r_1 \propto \sqrt{\frac{1}{M_1}}$$

$$r_2 \propto \sqrt{\frac{1}{d_2}} \quad or \quad r_2 \propto \sqrt{\frac{1}{M_2}}$$

Rearranging these relationships gives the following expression

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad or \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Where:

- **r** is the rate of diffusion, **d** is the density and **M** is the molecular mass of the gas.
- $\mathbf{r_1}$, $\mathbf{d_1}$ and $\mathbf{M_1}$ represent the rate of diffusion, density and molecular mass of gas 1.
- $\mathbf{r_2}$, $\mathbf{d_2}$ and $\mathbf{M_2}$ represent the rate of diffusion, density and molecular mass of

$$\frac{1}{2}m_A v_A^2 = \frac{1}{2}m_B v_B^2$$

$$m_A v_A^2 = m_B v_B^2$$

$$\frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\left(\frac{v_A}{v_B}\right)^2 = \frac{m_B}{m_A}$$

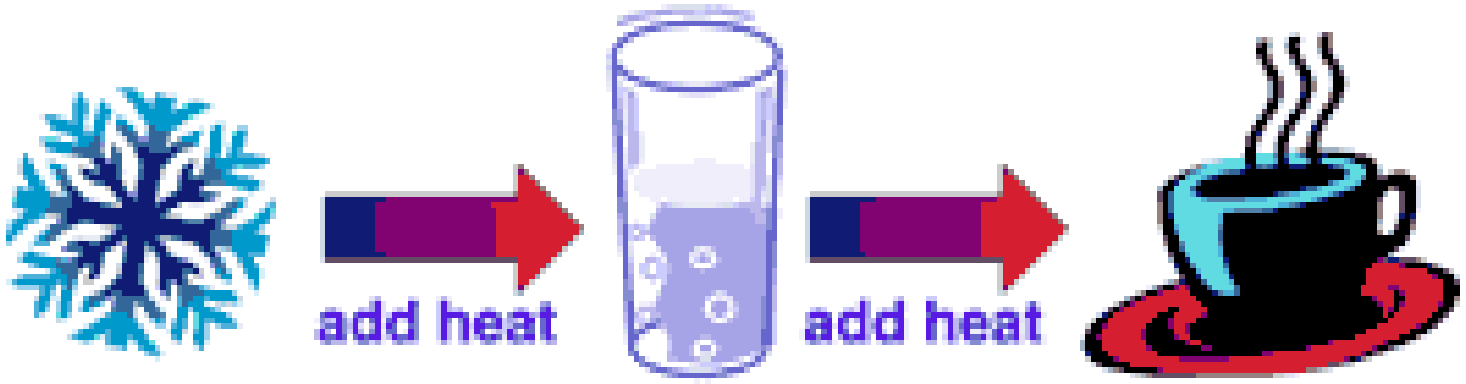
$$\frac{v_A}{v_B} = \sqrt{\frac{m_B}{m_A}}$$

$$\frac{effusion\,rate\,A}{effusion\,rate\,B} = \sqrt{\frac{molecular\,mass\,B}{molecular\,mass\,A}} \quad (Graham's\,Law)$$

3.5. Phase transition

What is a “phase transition” ?

- It is transformation of thermodynamic system from one phase to another or
- It is transformation of state of matter to another one by heat transfer



➢ Phases are states of matter characterized by distinct macroscopic properties

- Gas, liquid, and solid are *known as the three states of matter* or material, but each of solid and liquid states may exist in one or more forms.
- Thus, another term is required to describe the various forms, and the term phase is used.
- Each distinct form is called a phase, but the concept of phase defined as a homogeneous portion of a system, extends beyond a single material, because a phase may also involve several materials.

How Phase Transition works?

- There are ***two variables to consider*** when looking at phase transition, ***pressure (P) and temperature (T)***.
- For the gas state, The relationship between temperature and pressure is defined by the equations below:
- Ideal Gas Law: **PV=nRT**
- **van der Waals Equation of State:**

$$(P+a*n^2/V^2)(V-nb)=nRT$$

❖ **The ideal gas law assumes that no intermolecular forces are affecting the gas in any way, while the van der Waals equation includes two constants, a and b, that account for any intermolecular forces acting on the molecules of the gas.**

PHASE CHANGES/transition

- Depend on Temperature and Pressure

Affects:

  - Particle arrangement , Energy of particles and Distance between particles

- Phase change is either an exothermic or endothermic process
  - Exothermic – heat is given off i.e. Removal of E
  - Endothermic – absorption of heat i.e. Input of E
- Phase changes include:
  - Melting, evaporation/ boiling, condensation, freezing, sublimation and reverse sublimation

There are six ways a substance can change between these three phases; melting, freezing, evaporating, condensing, sublimation, and deposition .

These processes are reversible and each transfers between phases differently:

- ✓ ***Melting:*** The transition from the solid to the liquid phase
- ✓ ***Freezing*:** The transition from the liquid phase to the solid phase
- ✓ ***Evaporating:*** The transition from liquid phase to the gas phase
- ✓ ***Condensing:*** The transition from the gas phase to the liquid phase

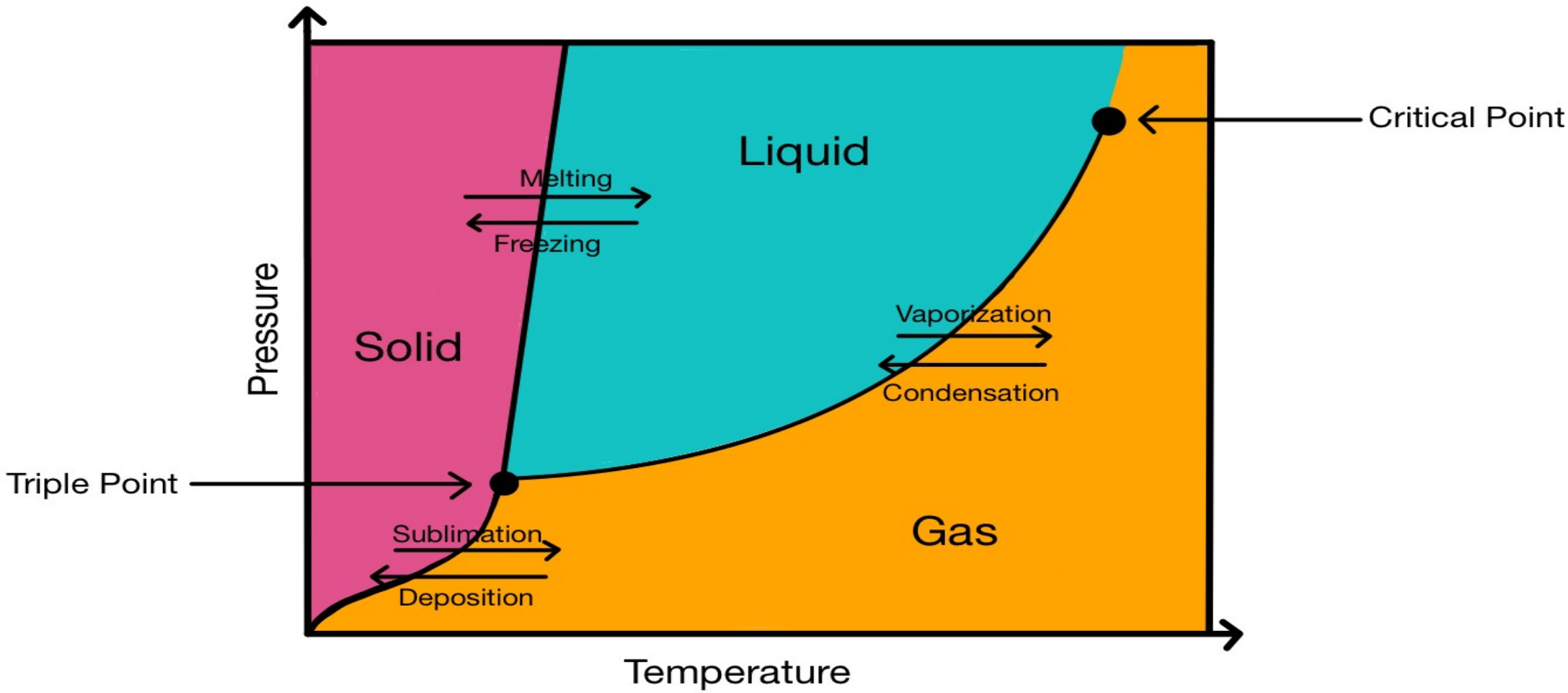
Phase Equilibrium

- **Component**: is either pure metal and/or compound of which an alloy is composed. They refer to the independent chemical species that comprise the system.
- **Solid Solution**: It consists of atoms of at least two different types where solute atoms occupy either substitutional or interstitial positions in the solvent lattice and the crystal structure of the solvent is maintained.
- **Solubility limit**: for almost all alloy systems, at specific temperature, a maximum of solute atoms can dissolve in solvent phase to form a solid solution. The limit is known as solubility limit.
- **Phase**: defined as a homogenous portion of a system that has uniform physical and chemical characteristics i.e. it is a physically distinct from other phases, chemically homogenous and mechanically separable portion of a system.

Equilibrium → state with minimum free energy under some specified combination of *temperature, pressure, and composition*

The driving force for a phase change is the minimization of free energy

Phase Diagram



GIBBS PHASE RULE

"On the Equilibrium of Heterogeneous Substances" 1874-78

$$P + F = C + 2$$

$$F = C - P + 2$$



Gibbs

P: number of **p**hases present at equilibrium
C: number of **c**omponents needed to describe the system
F: number of degrees of **f**reedom, e.g. T, P, composition

The *number of components* (C) is the minimum number of chemically independent constituents needed to describe the composition of the phases present in the system.

e.g.,
- salt water. $C = 2$ (NaCl and water)
- solid magnesium silicates. $C = 2$ (MgO and SiO_2)
- solid MgAl silicates. $C = 3$ (MgO, Al_2O_3, SiO_2)

The *degrees of freedom* (F) is the number of independent variables that must be specified to define the system completely.

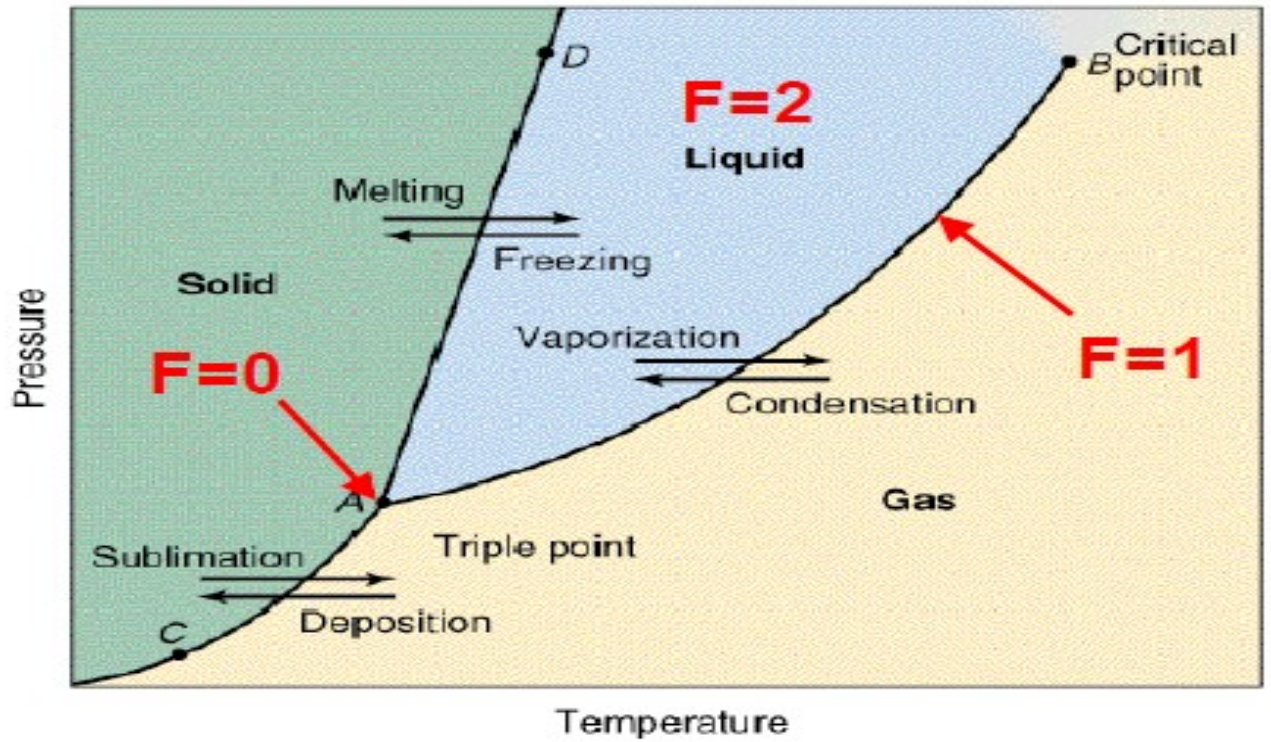
ONE COMPONENT PHASE DIAGRAMS

$$P + F = C + 2$$

with C = 1

$$\boxed{P + F = 3}$$

Composition is fixed, only T and P can vary



Three possibilities:

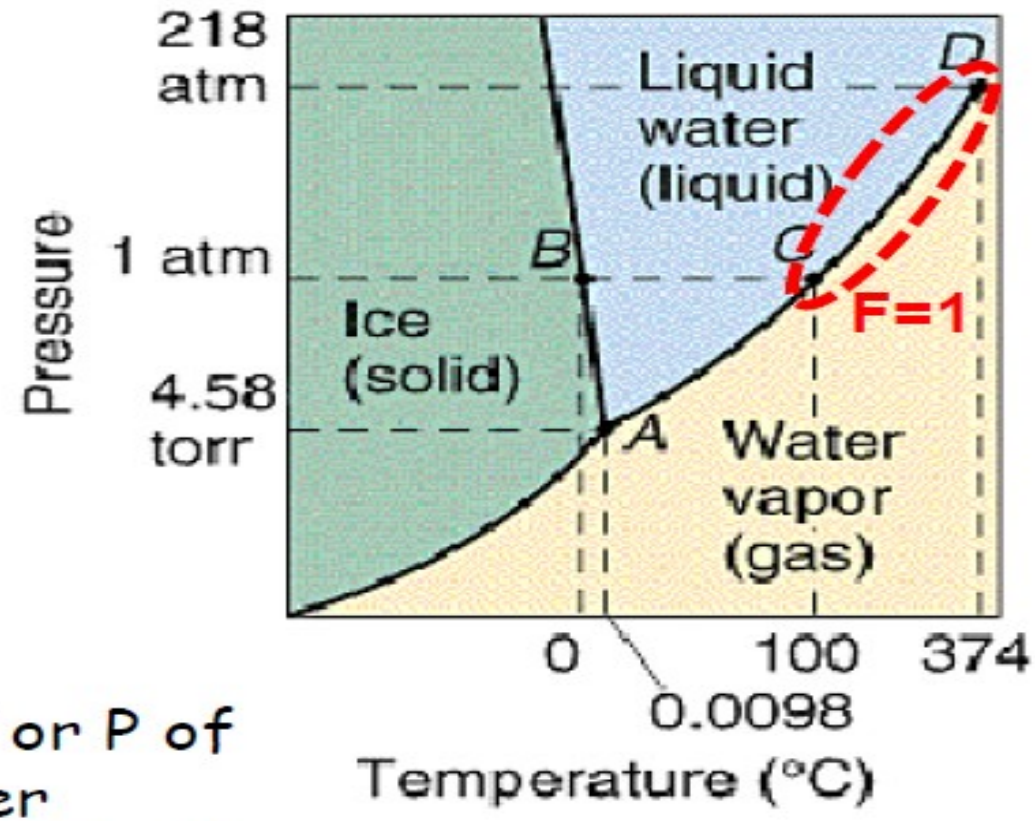
- P = 1 … F = 2 (bivariant phase field)
- P = 2 … F = 1 (univariant phase curve)
- P = 3 … F = 0 (invariant phase point)

EXAMPLE: BOILING WATER

$P + F = C + 2$

$C = 1$ (water)
$P = 2$ (vapor + liquid)

$F = 1$ (either T or P,
but not both)
$\rightarrow$ coexistence curve



*once we specify either T or P of our boiling water, the other variable is specified automatically

Exercise

1. During the process of freezing, what do you know about changes in the temperature and entropy of water?

Explanation:

The temperature of a sample during the process of a phase change will remain constant. Even though heat may be added or removed during this process, it is utilized to directly impact the formation and breaking of intermolecular interactions within the sample. The result is a change in the internal properties of the sample (the phase), with no actual change in temperature.

When a substance freezes it transitions from a liquid to a solid. The result is a much more ordered structure, as opposed to the fluidity associated with the liquid phase. Increasing order corresponds to a decrease in entropy because entropy is a measure of disorder.


Thank
you