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CHEMISTRY GRADE NINE
UNIT FIVE SUMMARY

ADDIS ABABA EDUCATION BUREAU

JUNE, 2020

UNIT 5
Physical States of Matter
Session 2

5. The Ideal Gas Equation

- An ideal gas is a hypothetical gas that obeys the gas laws. Real gases only obey the ideal gas laws closely at high temperature and low pressure.
- Under these conditions, their particles are very far apart.
- The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.
- We have considered the three laws that describe the behavior of gases as revealed by experimental observations:
- Boyle's law: $V \alpha \frac{1}{P}$ (at constant T and n)
- Charles' law: $V \alpha T$ (at constant P and n)
- Avogadro's law: $V \alpha n$ (at constant P and T)

- This relationship indicates how the volume of gas depends on pressure, temperature and number of moles.

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

or

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

- where ***R***, is a proportionality constant called the ***gas constant***.

$$PV = nRT \text{ (the ideal gas equation)}$$

- Thus, the ideal gas equation describes the relationship among the four variables ***P, V, T*** and ***n***.
- An ideal gas is a gas whose pressure-volume-temperature behavior can be completely explained by the ideal gas equation.

- At STP, the values of ***R*** can be calculated from the ideal gas equation.

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

$$= 0.082057\ \ \text{L.atm/K.mol}$$

$$= 8.314\ \ \ \text{L.kPa/K.mol}$$

$$= 8.314\ \text{J/mol.K}$$

- For calculations, we round off the value of R to three significant figures (0.0821 L.atm/K.mol) and use 22.4 L for the molar volume of a gas at STP.

Exercise

- Calculate the volume (in Liters) occupied by 4 moles of CO_2 at Temp= 273 K and Pressure= 1 atm (R= 0.082 L.atm/K.mol).

6. Graham's Law of Diffusion

- A gas tends to expand and occupy any space available to it. This spreading of gas molecules is called diffusion.

  How do you compare the rate of diffusion of molecules with different densities?

- Thomas Graham (1805 - 1869), an English chemist, studied the rate of diffusion of different gases.
- He found that gases having low densities diffuse faster than gases with high densities.
- He was able to describe quantitatively the relationship between the density of a gas and its rate of diffusion.
- 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:

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

- where r is the rate of diffusion, dis the density and M is the molecular mass of the gas. For two gases (Gas 1 and Gas 2), their rates of diffusion can be given as:

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

and

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

- Rearranging these relationships gives the following expression

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

- where
  - r_1, d_1 and M_1 represent the rate of diffusion, density and molecular mass of gas 1.
  - r_2, d_2 and M_2 represent the rate of diffusion, density and molecular mass of gas 2.

- Also the rate at which a gas diffuses is inversely proportional to the time taken. Mathematically,

$$r \propto \frac{1}{t}$$

- If two different gases (gas 1 and gas 2) under the same conditions of temperature and pressure diffuse through a porous container, then the time required to diffuse for the two gases can be given by the following formula:

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

- where t_1 and t_2 are the time taken, r_1 and r_2 are the rates, M_1 and M_2 are the molecular masses of Gas 1 and Gas 2 respectively.

Excercise

1. Which gas will diffuse faster,Hydrogen(H_2) or oxygen(O_2)?
2. What is the relative rate of diffusion for the above gases?

5.4 THE LIQUID STATE

- liquids have a definite volume and an indefinite shape.
- They take the shape of their containers to the level they fill. On the average, liquids are more dense than gases, but less dense than solids.
- As in a gas, particles in a liquid are in constant motion.
- However, the particles in a liquid are closer together than those in a gas.
- The attractive forces between particles in a liquid are more effective than between particles in a gas.
- This attraction between liquid particles is caused by the intermolecular forces (dipole-dipole forces, London dispersion forces, and hydrogen bonding).
- Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of liquid particles.
- Accordingly, liquid particles are not bound together in fixed positions.

Energy Changes in Liquids

- The process by which a liquid changes to a gas is known as vaporization.
- Evaporation is the process by which liquid molecules break freely from the liquid surface and enter the vapor phase.
- Evaporation is explained in terms of the energy possessed by the molecules on the surface of the liquid
- In an open container, evaporation continues until all of the liquid enters the vapor phase.
- However, liquids in a closed container behave differently.

- In closed containers, the vapor cannot escape. As the vapor concentration increases, some of the vapor molecules lose energy and return to the liquid state.
- When a vapor returns to the liquid state, it is said to condense.
- The process is called condensation. Evaporation and condensation are opposing processes.

Evaporation

Liquid  Gas

Condensation

- The rate of evaporation of a liquid depends on three factors. These are temperature, intermolecular forces, and surface area of the liquid.

- An increase in temperature increases the average kinetic energy of the molecules and this increases the tendency to change into the gaseous state.
- Some liquids evaporate readily at room temperature.
- Such liquids are said to be volatile.
- The volatile liquids have relatively weak forces of attraction between particles.
- Diethyl ether, ethyl alcohol, benzene and acetone are volatile liquids.
- Non-volatile liquids have a little tendency to evaporate at a given temperature.
- They have relatively stronger attractive forces between their molecules, e.g., sulphuric acid, water, and molten ionic compounds.

- ***Vapour pressure***: The partial pressure of the vapour above a liquid is called ***vapour pressure***.
- The vapour pressure of a liquid depends up on the temperature.
- At a given temperature, vapour pressure is constant. As the temperature increases, the vapour pressure of a liquid also increases due to high rate of evaporation.
- Vapour pressure depends also on the strength of the intermolecular forces between the particles of the liquid.
- The stronger the intermolecular forces, the lower the vapour pressure will be, because fewer particles will have enough kinetic energy to overcome the attractive force at a given temperature.
- For example, water and alcohol have relatively low vapour pressure.
- On the contrary, liquids with low intermolecular forces have high vapour pressures at room temperature.
- For example, diethyl ether, a nonpolar molecule with relatively weak dispersion forces, has a relatively higher vapour pressure.

Boiling and Boiling Point

What is the difference between evaporation and boiling? How does boiling point depend on the external pressure?

- Boiling is:
- the conversion of liquid to vapors within the liquid as well as at its surface.
- It occurs when the equilibrium vapour pressure of the liquid equals the atmospheric pressure.
- During evaporation only molecules at the surface escape into the vapour phase, but at the boiling point the molecules within the liquid have sufficient energy to overcome the intermolecular attractive forces of their neighbors, so bubbles of vapour are released at the surface.

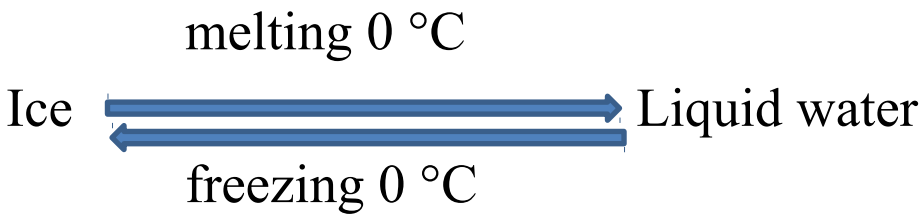
- If the temperature of the liquid is increased, the equilibrium vapour pressure also increases.
- Finally, the boiling point is reached.
- The boiling point of a liquid is the temperature at which the equilibrium vapour pressure of the liquid equals the atmospheric pressure.
- Therefore, the lower the atmospheric pressure, the lower the boiling point will be.
- If the external pressure is 1.0 atmosphere (760 mmHg), the boiling point is called normal boiling point.
- For instance water boils, at 100°C at 1.0 atmospheric pressure.
- Thus, the normal boiling point of water is 100°C. **Where will water boil first? In Addis Ababa or Hawassa? Why?**

- Boiling of a liquid requires a certain amount of heat energy to break the forces of attraction that holds the molecules together.
- The amount of heat energy necessary to bring about the vaporization of a fixed amount of a liquid at a fixed temperature to the gaseous state is called ***the heat of vaporization***.
- For example, the heat of vaporization per mole of water at 298 K and 1 atmosphere is 44.0 kJ.
- This is called ***the molar heat of vaporization (ΔHvap) of water***.
- Molar heat of vaporization ***(ΔHvap)*** and molar heat of condensation ***(ΔHcond)*** are equal in magnitude but opposite in sign, i.e., $\Delta H_{vap} = -\Delta H_{cond}$. Vaporization is an endothermic processes where as condensation is an exothermic process.

5.5 THE SOLID STATE

- When a solid is continuously heated its ordered crystalline structure is disturbed.
- The particles attain their freedom of motion gradually and melting (or fusion) takes place where the solid is converted into the liquid state.
- The temperature at which a crystalline solid is converted to a liquid is known as the melting point.
- In contrast when a liquid is cooled, the molecules become closer and closer, and the intermolecular forces of attraction become stronger and stronger.
- The particles arrange themselves into a regular pattern in the liquid.
- This process is called freezing or solidification.
- Note that the freezing point of a liquid is the same as the melting point of a crystalline solid. This means that:
    - Both melting and freezing take place at the same temperature.
    - Both liquid and solid phases coexist at equilibrium with each other at the melting or freezing point.

- To clarify this, let us consider ice, which melts at 0°C and water freezes at 0°C.
- Ice and water coexist in equilibrium with each other at 0°C as follows.



- The amount of heat needed to convert one gram of a solid to a liquid at the melting point is called heat of fusion.
- The molar heat of fusion or molar enthalpy of fusion ***(ΔHfus)*** is the quantity of heat needed to convert one mole of a solid at its melting point to the liquid state.
- For example, the molar heat of fusion of ice is 6.01 kJ at 0°C.
- This is the amount of energy needed to break the attractive forces in the solid at the melting point.
- Melting requires the supply of energy; therefore, it is an endothermic process.

- During the process of solidification, an amount of heat equal to the heat of fusion must be liberated.
- This quantity of heat liberated, which is exactly equal to the heat of fusion, is called the heat of solidification or heat of crystallization.
- The molar heat of crystallization (ΔHcryst) is the quantity of energy that must be removed from one mole of a liquid to convert it to the solid state at its freezing point.
- $\Delta H_{cryst} = -\Delta H_{fus}$ Although the motion of the particles in a solid are more restricted than those in a liquid, many solids have a significant vapour pressure and evaporate directly from the solid to the vapour state without passing through the liquid state.
- This process is called sublimation



Solid Vapour(gas)

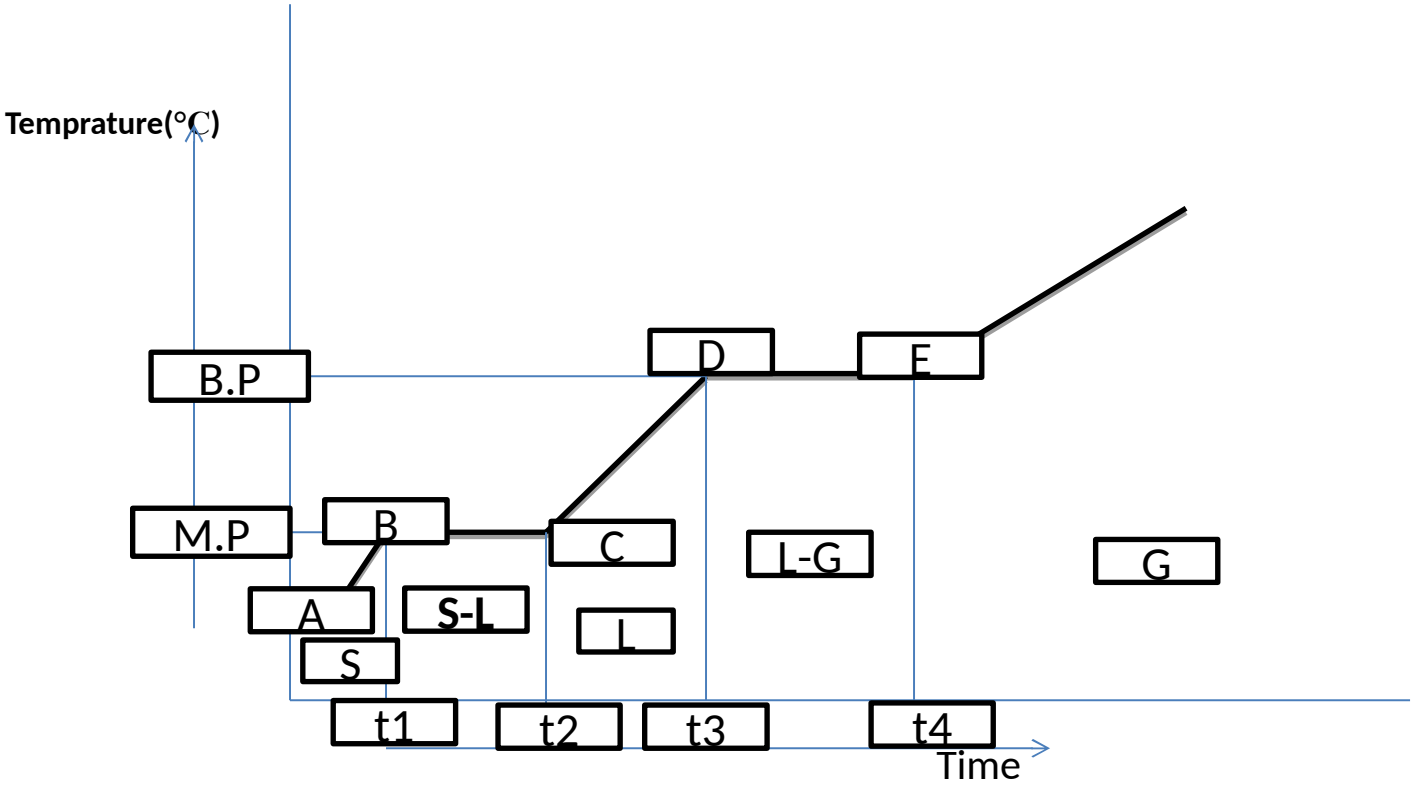
- During the sublimation process heat energy is absorbed.
- That is, it is an endothermic process. Molar heat of sublimation ***(ΔHfus)*** is the quantity of heat required to convert one mole of a solid to a gas at its sublimation point.
- The heat (enthalpy) of sublimation is related to the enthalpies of fusion and vaporization by:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

Phase Changes and Energy Changes in Solids

- A phase is any part of a system that has uniform composition and properties.
- A state of matter represents a phase.
- Most solid substances undergo two changes of state when heated.
- A solid change to a liquid at the melting point, and the liquid changes to vapour at the boiling point.
- To understand state changes, we will consider the heating curve for a substance.
- A heating curve is a plot of temperature verses the uniform addition of heat.
- This can be illustrated for a hypothetical substance, in which the temperature of the substance is on the vertical axis and the passage of time during which heat is added to the substance is on the horizontal axis.
- The changes in the temperature and phases of a pure substance as it is heated, beginning with a solid and continuing to the gaseous state as described.

Heating Curve



- Initially, the substance exists in the solid state, and the addition of heat increases its temperature.
- When the solid is heated, its temperature rises (A to B) until it reaches the melting point (point B), and the temperature remains constant (B to C) until all the solid is converted to a liquid (point C).
- The added heat energy is used to break the intermolecular forces, thus disrupting the solid structure.
- At point C phase change is completed.
- Once melting is completed, heating of the liquid raises its temperature (C to D) until the boiling point is reached at point D.
- In region (D to E) the addition of heat is utilized to break the intermolecular forces of the liquid to change it to a gas.
- When all the liquid has been converted into a gas, addition of heat simply raises the temperature of the gas.