GENERAL CHEMISTRY 2

Quarter 1

STEM | LABADIA - TUGA - GUNDAY - BENIGAY | SEM 1 2022

LESSON 1: INTERMOLECULAR FORCES OF LIQUIDS AND SOLIDS

KINETIC MOLECULAR THEORY

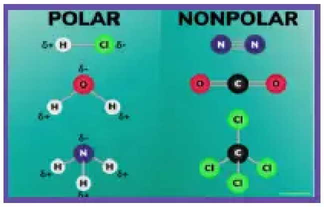
1. All particles are composed of very small particles called atoms or molecules.
2. The distances between gas particles are so great that the volume of the gas is mostly empty space.
3. There are no repulsive or attractive forces between gas particles.
4. These particles are in constant random motion.
5. The kinetic energy of the particles depends on the temperature of the gas.

A Molecule will be Nonpolar if:

- All of the terminal atoms (or groups) are the same.
- All of the terminal atoms (or groups) are symmetrically arranged around the central atom.
- The terminal atoms (or groups) have the same charges.

A Molecule will be Polar if:

- One or more terminal atoms differ from each other
- At least one polar bond is present
- The terminal atoms are not symmetrically arranged
- The molecule has one slightly positive end one slightly negative end



PROPERTIES OF MATTER

Properties of matter	Molecular Behavior		
	Gas	Liquid	Solid
Volume/shape	Assumes volume and shape of the container	Fixed volume, assumes shape of occupied part of container	Fixed volume, fixed (regardless the shape and size of the container)
Density	Low	High	High
Compressibility	Easy to compress	Cannot be appreciably compressed	Cannot be appreciate compressed
Motion of molecules	Random, fast, cover large distances	Random, medium speed, limited distance	Vibration in place

- An increase in temperature results in increased kinetic energies of gases dissolving in liquid.
- This increased motion enables the dissolved gas to break intermolecular forces with the solvent, and escape the solution.

Example:

A warm bottle of carbonated drink does not taste as good as a cold one, because there is less carbon dioxide dissolved in the warm bottle.

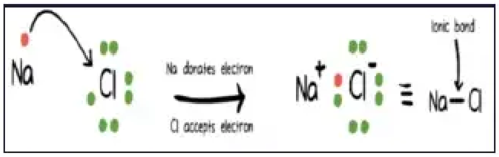
INTRAMOLECULAR FORCES

- Attractive forces that act **within** molecules or particles in the solid or liquid states.
- Are relatively strong because their charges are larger and closer together.

TYPES OF INTRAMOLECULAR FORCES OF ATTRACTION:

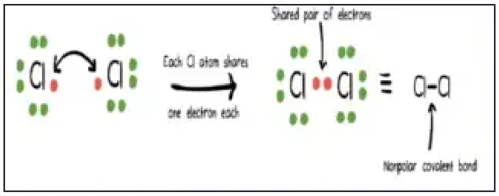
1. Ionic bond

- occurs between metal and nonmetal compounds.



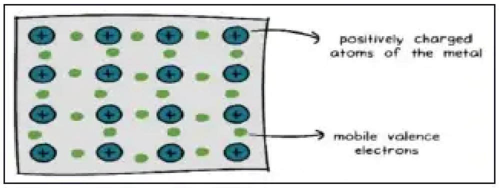
2. Covalent bond

- occurs between two nonmetal compounds.



3. Metallic bond

- occurs between metal atoms.



INTERMOLECULAR FORCES

- Attractive forces that act **between** molecules or particles in the solid or liquid states. (much weaker than bonding forces)
- Occur between neighboring molecules as a result of partial charges or between ions and molecules.
- They are relatively weak because they involve smaller charges that are farther apart.
- Intermolecular forces as a whole are usually called vander Waals forces after the Dutch scientiest Johannes van der Waals (1837-1923).

Why melting points and boiling points of substances can be used as indicators of strength of intermolecular forces operating in given solids and liquids?

- When a solid melts, or a liquid boils, the particles move away from each other.
- As they do, intermolecular forces of attraction are broken.

TYPES OF INTERMOLECULAR FORCES

1. Ion-Dipole Forces

- Results when an ion and the partial charge on the end of polar molecule attract each other

2. Dipole-Dipole Forces

- Attractive forces between polar molecules.
- In polar molecules the electrons are unevenly distributed because some elements are more electronegative than the others.

3. Hydrogen bond

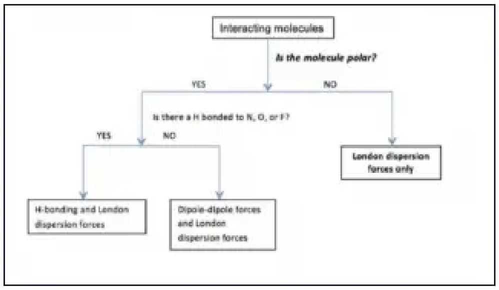
- A special type of dipole-dipole interaction between the hydrogen atom in a polar bond, such as N-H, O-H, or F-H, and an electronegative O, N, F atom. Hydrogen bonds between water molecules are particularly strong.

4. Dispersion Forces

- Forces of attraction result from the temporary dipole moments induced in ordinarily nonpolar molecules.
- These forces are present between ALL types of molecules due to the movement of electrons.
- As electrons move around the nucleus, the uneven distribution causes momentarily charge separations.
- Slightly positive sides of the molecule are attracted to the slightly negative side of the adjacent molecule.

5. Polarizability

- Refers to the ease with which the electron distribution can be distorted.
- The larger number of electrons, the greater its Polarizability.
- Thus dispersion forces may be the weakest of intermolecular forces that can exist between two molecules, but the larger the atoms present, the stronger the dispersion forces.



LESSON 2: PROPERTIES OF LIQUIDS

Fluid

- a gas or a liquid; substance that can flow.

Surface tension

- the measure of the elastic force on the surface of a liquid. It is the amount of energy required to stretch or increase the surface of a liquid by a unit area.

Capillary action

- the tendency of a liquid to rise in narrow tubes makes it easily drawn into small openings.

Viscosity

- a measure of a fluid's resistance to flow.

Vapor

- a gaseous substance that exist naturally as a liquid or solid at normal temperature.

Vaporization

- The change of phase from liquid to vapor (gaseous phase).

Vapor pressure of a liquid

- The equilibrium pressure of a vapor above its liquid; that is, the pressure exerted by the vapor above the surface of the liquid in a closed container.

Boiling point

- The temperature at which a liquid boils. The boiling point of a liquid when the external pressure is 1 atm is called the normal boiling point.

Molar heat of vaporization (ΔH_{vap})

- The energy (usually in kilojoules) required to vaporize 1 mole of a liquid at a given temperature.

Surface tension

- is the measure of the elastic force on the surface of a liquid. It is the amount of energy required to stretch or increase the surface of a liquid by a unit area.
- It is manifested as some sort of skin on the surface of a liquid or in a drop of liquid.
- Molecules within a liquid are pulled in all directions by intermolecular forces.
- Molecules at the surface are pulled downward and sideways by other molecules, not upward away from the surface.
- These intermolecular forces tend to pull the molecules into the liquid and cause the surface to tighten like an elastic film or "skin"
- Liquids that have strong intermolecular forces also have high surface tension.

ADDITIONAL INFORMATION:

Why will water form a spherical droplet on a sheet of plastic, but kerosene will spread?

- The water forms a spherical droplet on plastic since it consists of strong intermolecular forces and higher surface tension compared to kerosene.

CAPILLARY ACTION

- is the tendency of a liquid to rise in narrow tubes or be drawn into small openings such as those between grains of a rock.
- Capillary action, also known as capillarity, is a result of intermolecular attraction between the liquid and solid materials.

TWO TYPES OF FORCES INVOLVED IN CAPILLARY ACTION:

1. Cohesion

- is the the intermolecular attraction between like molecules (the liquid molecules)

2. Adhesion

- Is an attraction between unlike molecules (such as those in water and in the particles that make up the glass tube).

These forces also define the shape of the surface of a liquid in a cylindrical container (the **meniscus**!)

Mercury & Water In A Test Tube

ADDITIONAL INFORMATION:

Concave Meniscus
- The molecules of the liquid are strongly attracted to each other.
- Cohesion force of the liquid is great,

When the cohesive forces between the liquid molecules are lesser than the adhesive forces between the liquid and the walls of the container, the surface of the liquid is concave.
Example: water in a glass container

Convex Meniscus
- The adhesive force is great within the molecules of the container and liquid.

When both adhesive and cohesive forces are equal, the surface is horizontal.
Ex. distilled water in a silver vessel

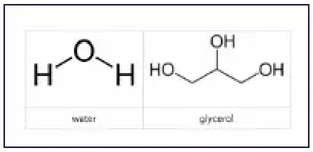
Meniscus
- surface shape of a liquid

Viscosity
- is a measure of a fluid's resistance to flow. The greater the viscosity, the slower the liquid flows.
- is expressed in units of centipoise.

Substances with lower viscosities include carbon tetrachloride and benzene. Glycerol has a resistance to flow of more than a thousand times greater than water.

VISCOSITY OF LIQUIDS OF SOME PURE SUBSTANCES

LIQUIDS	VISCOSITY (IN NS/M2) AT 20OC
Acetone (C3H6O)	3.16 x 10-4
Benzene (C6H6)	6.25 x 10-4
Carbon tetrachloride (CCl4)	9.69 x 10-4
Diethyl ether (C2H5OC2H5)	2.33 x 10-4
Ethanol (C2H5OH)	1.20 x 10-3
Glycerol (C3H8O3)	1.49
Mercury (Hg)	1.55 x 10-3
Water (H2O)	1.01 x 10-3



Given molecular structures of water and glycerol, can you tell why glycerol has a higher viscosity than water?
- The larger number of –OH groups allow glycerol to form more H-bonds with other glycerol molecules, making its intermolecular forces stronger than those of water, and its resistance to flow greater.

Consider the table of viscosities that follow. All the substances in the list are hydrocarbons and nonpolar. What causes the differences in viscosities of the Hydrocarbons in the list?

- The size of the molecules. The larger the molecule, even if it is nonpolar, the stronger the intermolecular forces and the greater the viscosity compared to nonpolar substances made up of higher molecules.

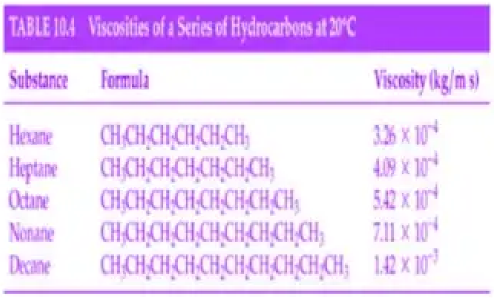


TABLE 10.4 Viscosities of a Series of Hydrocarbons at 20°C

Substance	Formula	Viscosity (kg/m s)
Hexane	$CH_3CH_2CH_2CH_2CH_2CH_3$	3.26×10^{-4}
Heptane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$	4.09×10^{-4}
Octane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$	5.42×10^{-4}
Nonane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$	7.11×10^{-4}
Decane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$	1.42×10^{-3}

- Liquids that have strong intermolecular forces have higher viscosities than those that have weak intermolecular forces.

What effect temperature would have on viscosity?

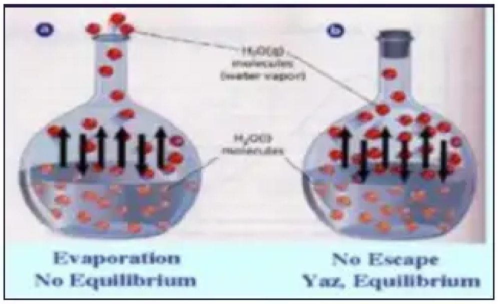
- Viscosity decreases as temperature increases: hot molasses flows much faster than cold molasses. The viscosities of some familiar liquids in the table below were measured at 20 OC, except for lava (ranges between 700 to 1200 °C.

LIQUIDS	VISCOSITY (IN NS/M2) AT 20OC
Water	1
Milk	3
Blood	4 to 10
Castor oil	1000
Latex house paint	1500
Hotcake syrup	5000
Honey	10,000
Hershey's chocolate syrup	10,000 to 25,000
Ketchup	50,000
Peanut butter	250,000
Lava	≈ 4,300,000

- The water molecules in the liquid evaporate and go into the vapor phase. In the open flask, some of the water molecules in the vapor phase find their way out of the flask and are lost to the atmosphere.

Vapor Pressure of a Liquid

What is happening to the water molecules in the two flasks shown in the picture?



- The water molecules in the liquid evaporate and go into the vapor phase. In the open flask, some of the water molecules in the vapor phase find their way out of the flask and are lost to the atmosphere.
- When a liquid evaporates to a gas in a closed container, the molecules cannot escape.
- Gas molecules move in random directions, collide with other gas particles and the walls of the container
- Some will strike the liquid surface and condense back into it. In the closed flask, none of the gas particles are able to get out of the

container, and eventually, the number of molecules that go into the gaseous state would equal the number of molecules that condense back.

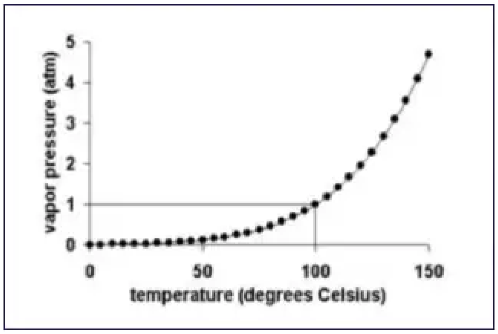
- When the rate of condensation of the gas becomes equal to the rate of evaporation of the liquid, the gas in the container is said to be in equilibrium with the liquid.

liquid $\rightleftharpoons$ vapor (gas)

Vapor Pressure

- Like any gas sample, the molecules in the gaseous state over its liquid create a pressure. The greater the number of gaseous particles, the greater the pressure exerted by the gas. The pressure exerted by the gas in equilibrium with a liquid in a closed container at a given temperature is called the **equilibrium vapor pressure or simply vapor pressure of the liquid.**
- The **equilibrium vapor pressure** is the maximum vapor pressure of a liquid at a given temperature and that it is constant at a constant temperature. It increases with temperature.
- **Vapor pressure** is independent of the amount of liquid as well as the surface area of the liquid in contact with the gas.

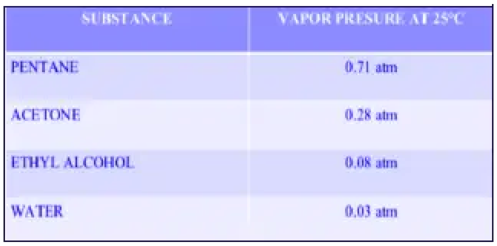
RELATIONSHIP OF VAPOR PRESSURE TO TEMPERATURE



- As the temperature increases, the vapor pressure of water also increases.

When temperature is high, more molecules have enough energy to escape from the liquid. At a lower temperature, fewer molecules have sufficient energy to escape from the liquid.

RELATION OF VAPOR PRESSURE TO STRENGTH OF INTERMOLECULAR FORCES



SUBSTANCE	VAPOR PRESURE AT 25°C
PENTANE	0.71 atm
ACETONE	0.28 atm
ETHYL ALCOHOL	0.08 atm
WATER	0.03 atm

- Ethyl alcohol and water have very low vapor pressures. Both liquids have the strong dipole-dipole interaction called hydrogen bonding. Acetone is polar but does not have H-bonding. Its vapor pressure is of intermediate value. Pentane is a nonpolar substance, and its vapor pressure is high compared to those of water and ethyl alcohol.

- When liquids evaporate, the molecules have to have sufficient energy to break the attractive forces that hold them in the liquid state. The stronger these intermolecular forces are, the greater the amount of energy needed to break them. The stronger the intermolecular forces of attraction, the lower the vapor pressure of a liquid.
- The stronger the intermolecular forces of attraction, the lower the vapor pressure of a liquid.

Molar Heat of Vaporization

- The molar heat of vaporization (ΔHvap) is the energy required to vaporize 1 mole of a liquid at a given temperature. H is the symbol for enthalpy, which means heat content at a given standard condition.

SUBSTANCE	ΔHVAP (KJ/ MOL)	BOILING POINT* (OC)
Argon (Ar)12)	6.3	-186
Pentane(C_5H	26.5	36.1
Acetone (CH_3COCH_3)	30.3	56.5
Ethanol (C_2H_5OH)	39.3	78.3
Water (H_2O)	40.79	100

The relationship between vapor pressure and strength of intermolecular forces is consistent with the trends in two other properties of liquids, the enthalpy or molar heat of vaporization, and the boiling point of the liquid.

Relate Molar Heat of Vaporization To Strength Of Intermolecular Forces

- The heat of vaporization may be considered a measure of the strength of intermolecular forces in a liquid. If the intermolecular attraction is strong, it takes a lot of energy to free the molecules from the liquid phase and the heat of vaporization will be high.
- It is easier to vaporize acetone (lower Hvap) than water (higher Hvap) at a given temperature, and more acetone escapes into the vapor phase at a given temperature. Acetone is a polar substance but has no H-bonding. It has weaker intermolecular forces than water, and therefore acetone molecules are held less tightly to one another in the liquid phase.

Boiling Point

- The **boiling point** is the temperature at which the vapor pressure of a liquid is equal to the external pressure.
- The **normal boiling point** is the temperature at which the liquid converts to a gas when the external pressure is **1 atm**. The normal boiling point of water is **100°C.**

Relate boiling point to molar heat of vaporization.

- The higher ΔHvap (Molar heat of vaporization) , the higher the boiling point.
- The boiling points of substances often reflect the strength of the intermolecular forces operating

among the molecules. At the BP, enough energy must be supplied to overcome the attractive forces among molecules before they can enter the vapor phase.

LESSON 3: PROPERTIES OF SOLIDS

Fluid
- A gas or a liquid; a substance that can flow.

Crystal or crystalline solid
- A solid material whose components, such as atoms, molecules or ions, are arranged in a highly ordered microscopic structure.

Ion
- An atom or group of atoms that has a net positive or negative charge.

Ionic crystal
- A solid that consists of positively and negatively charged ions held together by electrostatic forces.

Electrostatic bonding
- The attraction between oppositely charged ions in a chemical compound.

Ionic bond
- The electrostatic force that holds ions together in an ionic compound.

Network solid or covalent network crystal
- A solid that may be a chemical compound (or element) in which atoms are bonded by covalent bonds in a continuous network extending throughout the material.

Molecular crystal
- A solid composed of molecules held together by van der Waals forces (dispersion force, dipole-dipole attraction, hydrogen bonding).

Covalent bond
- A bond in which one or more pairs of electrons are shared by two atoms.

Dispersion forces
- Interactions that are the result from temporary dipole moments induced in ordinarily nonpolar molecules.

Dipole-dipole forces
- Attractive forces between polar molecules (molecules that possess permanent dipole moments).

Hydrogen bond
- A special type of dipole-dipole interaction between the hydrogen atom in a polar bond such as N–H, O–H, or F–H, and any of the electronegative atoms O, N, or F.

Crystal lattice
- The regular repeating structure of a crystalline solid.

Unit cell
- The smallest subunit of a crystal lattice that can be repeated over and over to make the entire crystal.

Gemstone
- may be a pure chemical element
- a relatively simple chemical compound (quartz is silicon dioxide, SiO2)
- a more complex mixture of various compounds and elements
- The great majority of familiar gem materials are oxides or silicates (e.g., they contain oxygen and perhaps silicon)
- formed as crystals during the cooling of the earth's crust over past millennia.
- may be formed in single or multiple discrete crystals (such as diamond).

In general, **larger crystals** were formed in areas of slow cooling for molten rock, and smaller crystals in areas of more rapid Cooling.

TWO GENERAL TYPES OF SOLIDS

- Solids can be categorized into two groups: the **crystalline solids** and the **amorphous solids**.

FEATURES THAT CAN BE USED TO DISTINGUISH A CRYSTALLINE SOLID FROM AN AMORPHOUS SOLID:

1. Arrangement of Particles

a. Crystalline Solids

- They can form a regular repeating three-dimensional structure called a crystal lattice (the regular repeating structure of a crystalline solid).
- are arranged in fixed geometric patterns or lattices. Examples of crystalline solids are ice and sodium chloride (NaCl), copper sulfate (CuSO4), diamond, graphite, and sugar (C12H22O11).
- The ordered arrangement of their units maximizes the space they occupy and are essentially incompressible.

b. Amorphous Solids

- They can aggregate with no particular long range order
- Amorphous solids have a random orientation of particles. Examples of amorphous solids are glass, plastic, coal, and rubber.
- They are considered super-cooled liquids where molecules are arranged in a random manner similar to the liquid state.

2. Behaviour when heated

a. Crystalline Solids

- become liquids at a specific temperature (i.e. the melting point).
- At this temperature, physical properties of the crystalline solids change sharply.

b. Amorphous Solids

- soften gradually when they are heated. They tend to melt over a wide range of temperatures. This behavior is a result of the variation in the arrangement of particles in their structures, causing some parts of the solid to melt ahead of other parts.

What is the distinguishing feature of crystalline solids?

- Crystalline solids are characterized by a regular repeating structure called the **crystal lattice**.

How are the structures of crystals determined?:

- **X-Ray Diffraction** is a technique used to determine the atomic and molecular structure of a crystal, wherein atoms cause beams of incident X-rays to diffract into many specific directions.

Sugar

- Cane stalks are shredded and squeezed to extract its natural juice and boiled until it thickens and molasses – rich sugar crystals begin to settle and are sent to a centrifuge to remove molasses.

Amethyst

- formed in silica-rich liquids deposited in gas cavities of lava that occur in crystalline masses.

Uric Acid

- Byproduct of protein, digestion, and among healthy individuals and is removed from the bloodstream and excreted by the kidney. The excess uric acid is known as gout

Snow

- Formed when temperatures are low and there is moisture in the form of tiny ice crystals in the atmosphere. The tiny ice crystals stick together in clouds to form snowflakes.

Crystallization

- refers to the formation of solid crystals from a homogeneous solution. It is essentially a solid-liquid separation technique and a very important one at that.

4 TYPES OF CRYSTALS:

1. Metallic crystals

- are made of atoms that readily lose electrons to form positive ions (cations), but no atoms in the crystal would readily gain electrons. The metal atoms give up their electrons to the whole crystal, creating a structure made up of an orderly arrangement of cations surrounded by delocalized electrons that move around the crystal. The crystal is held together by electrostatic interactions between the cations and delocalized electrons. These interactions are called **metallic bonds.** This model of metallic bonding is called the **"sea of electrons" model.**

OBSERVED PROPERTY	INFERENCE ABOUT THE STRUCTURE
Dense	Atoms are packed close together.
High Melting Point	Strong attractive forces hold the atoms in the crystal.
Good electrical conductor	Charged particles move through the crystal.
Good heat conductor	Particles can move through the crystal.
Malleable and ductile	When the crystal is deformed or stress is applied, the attractive forces are not broken.
Lustrous	Light is easily absorbed and emitted back.

PROPERTIES OF METALS/METALLIC CRYSTALS

- **High melting point** - large amount of energy is needed to melt crystal since the forces of attraction to be broken are numerous and extend throughout crystal
- **Dense** - atoms are packed together
- **Electric conductivity** - delocalized electrons move throughout the crystal
- **Thermal/ heat conductor** - decolorized electrons collide with each other as they move throughout the crystal and through these collisions that kinetic energy is transferred.
- **Malleability/ ductility** - when stress is applied to metal, metal cations shift in position, but the mobile electrons follow the movement of cations. Attractive forces between cations and mobile electrons are not broken.
- **Luster** - motion and collision of electrons allow it to gain and lose energy.

2. Ionic crystals

- Ionic crystals are made of ions (cations and anions). These ions form strong electrostatic interactions that hold the crystal lattice together. The electrostatic attractions are numerous and extend throughout the crystal since each ion is surrounded by several ions of opposite charge, making ionic crystals hard and of high melting points.

OBSERVED PROPERTY	INFERENCE ABOUT THE STRUCTURE
Hard	strong attractive forces hold the crystal together
High melting point	strong attractive forces have to be broken melt the crystal
Poor electrical conductor in the solid state	no charged particles move through the crystals
Good electrical conductor in molten state	mobile charged particles are present in molten state
Brittle	deformation or shift of particles cause attractive forces to be broken

- Ionic substances can conduct electricity in the liquid or molten state or when dissolved in water, indicating that in these states, charged particles are able to move and carry electricity. However, the solid state is generally nonconducting since the ions are in fixed positions in the crystal lattice and are unable to move from one point to another.
- Ionic crystals are brittle, and would shatter into small pieces when deformed or when pressure is applied on the crystal. The shifting of ions causes repulsions between particles of like charges.

3. Molecular crystals

- Molecular crystals are made of atoms, such as in noble gases, or molecules, such as in sugar, $C12H22O11$, iodine, $I2$, and naphthalene, $C10H8$. The atoms or molecules are held together by a mix of hydrogen bonding/dipole-dipole and dispersion forces, and these are the attractive forces that are broken when the crystal melts. Hence, most molecular crystals have relatively low melting points.

OBSERVED PROPERTY	INFERENCE ABOUT THE STRUCTURE
Soft	Weak attractive forces hold the crystal together.
Low melting point	Weak attractive forces are broken when crystals melt
Poor electrical conductor in the solid and molten states	No charged particles move through the crystal
Poor heat conductor	No particles can move easily throughout the crystal.
Brittle	Deformation or shift of particles cause

	attractive forces to be broken.

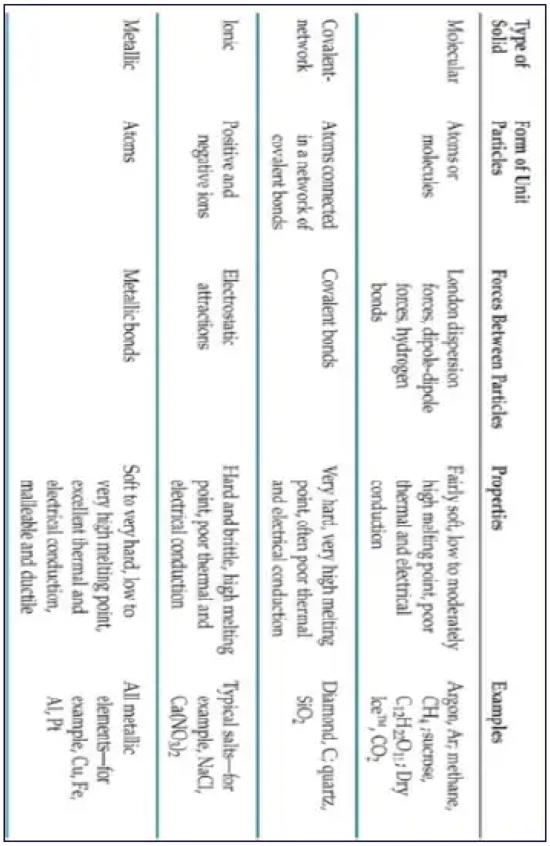
- The valence electrons of molecular substances are used in bonding, and cannot move about the crystal structure. Hence, the crystals are nonconducting. The absence of any mobile particles make molecular crystals unable to transmit heat fast. The crystals are brittle because the attractive forces that hold the molecules in the crystal are highly directional and a shift in positions of the molecules would break them.

4. Covalent network crystals

- crystals are made of atoms in which each atom is covalently bonded to its nearest neighbors. The atoms can be made of one type of atom (e.g. Cdiamond and Cgraphite) or can be made of different atoms (e.g. SiO2 and BN). In a network solid, there are no individual molecules and the entire crystal may be considered one very large molecule.

OBSERVED PROPERTY	INFERENCE ABOUT THE STRUCTURE
Hard	Strong attractive forces hold the crystal together.
Very high melting point	Strong attractive forces have to be broken in order to melt crystals
Poor electrical conductor in the solid and molten states	No charged particles move through the crystal
Poor heat conductor	No particles can move easily throughout the crystal.
Brittle	Deformation or shift of particles cause attractive forces to be broken.

- Diamonds are an example of a covalent network solid in which atoms are covalently bonded with each other. They tend to be hard and have high melting points.



Type of Solid	Form of Unit Particles	Forces Between Particles	Properties	Examples
Molecular	Atoms or molecules	London dispersion forces, dipole-dipole forces, hydrogen bonds	Fairly soft, low to moderately high melting point, poor thermal and electrical conduction	Argon, Ar, methane, CH_4, sucrose, $C_{12}H_{22}O_{11}$; Dry Ice™, CO_2
Covalent-network	Atoms connected in a network of covalent bonds	Covalent bonds	Very hard, very high melting point, often poor thermal and electrical conduction	Diamond, C; quartz, SiO_2
Ionic	Positive and negative ions	Electrostatic attractions	Hard and brittle, high melting point, poor thermal and electrical conduction	Typical salts—for example, NaCl, $Ca(NO_3)_2$
Metallic	Atoms	Metallic bonds	Soft to very hard, low to very high melting point, excellent thermal and electrical conduction, malleable and ductile	All metallic elements—for example, Cu, Fe, Al, Pt

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Quarter 1

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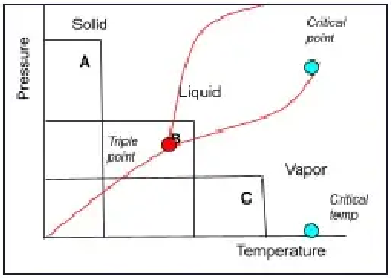
LESSON 4: INTERMOLECULAR FORCES OF LIQUIDS & SOLIDS

PHASE DIAGRAM

- graphical representation of physical states of substance of physical states of substance under different conditions of temperature and pressure
- gives possible combinations of pressure and temperature at which physical state a substance would be observed

FEATURES OF PHASE DIAGRAM:

- plots of pressure vs. celsius
- divided into 3: solid, liquid, gaseous
- boundary between liquid and gaseous regions stops at critical temperature for substance



-boundary between liquid and gaseous regions at point C is critical temperature

Freezing (Melting) Curve

- Represents the transition between liquid and solid states

Vaporization (Condensation) Curve

- Represents the transition between gaseous and liquid states

Sublimation (Deposition) Curve

- Represents the transition between gaseous and solid states

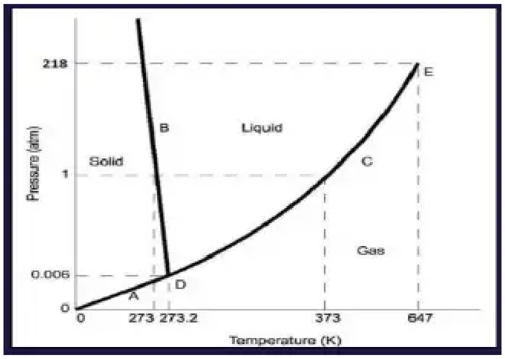
Triple point

- Combinations of pressure and temperature at which all 3 phases of matter area at equilibrium
- 3 states of matter coexist
- Represents of conditions of solid-liquid, liquid-vapor, and solid-vapor equilibrium meet at triple point

Critical point

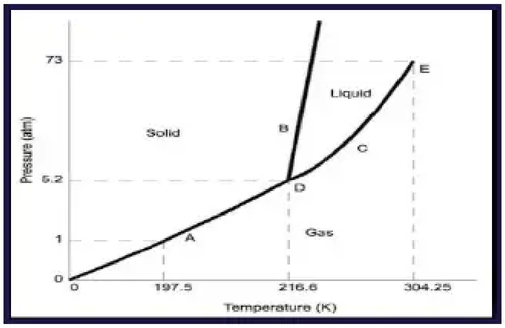
- Terminates liquid/gas phase line
- Set of temperature and pressure on phase diagram where liquid and gaseous phases of substance merge together into single phase
- Beyond temperature of critical point, merged single phase is **supercritical fluid**

PHASE DIAGRAM FOR WATER



- Triple point: 0.006 atm and 273. 2 K
- Critical temperature: 647 K (374°C)
- Critical pressure: 218 atm
- It is impossible to convert water from gas to liquid by compressing it above this temperature

PHASE DIAGRAM FOR CARBON DIOXIDE



- solid and carbon dioxide is known as "dry ice"
- at 1 atm, CO2 will sublime at temperature of 197.5 K (-75.5 °C)
- there is no liquid CO2 under normal conditions – only the solid or the vapor

LESSON 5: CONCENTRATION OF SOLUTIONS

TYPES OF SOLUTIONS:

a. **Dilute Solutions**
    - contains a small amount of solute dissolved compared to the solvent.

b. **Concentrated Solutions**
    - contains a large amount of solute compared to the amount of solvent.

c. **Saturated Solutions**
    - contains the maximum amount of a solute that will dissolve in a given solvent at a specific temperature.

d. **Unsaturated Solutions**
    - contains less solute than it has the capacity to dissolve.

e. **Supersaturated Solutions**
    - contains more solute than is present in a saturated solution.

PERCENT BY MASS AND PERCENT BY VOLUME

- It is defined as the quotient between the mass of the solute and the volume of solution multiplied by 100%.
- It is defined as the quotient between the mass of the solute and the volume of solution multiplied by 100%.

FORMULA:

1. **Percent by mass = Mass of Solute/Mass of Solution x 100%**
2. **Percent by volume= Volume of Solute/Volume of Solution x 100%**

PARTS PER MILLION (ppm)

- Is a unit for expressing every dilute concentration.
- It is commonly used to express the concentrations of pollutants in air or in water. Components of gas mixtures present in very small amounts are usually expressed in ppm by volume as defined by the equation.

FORMULA:

1. **ppm of component = Volume of Component/Total Volume of Solution x 10^6 ppm**

MOLARITY

- defined as the number of moles of solute per liter of solution.
- UNIT: mol/L or M

FORMULA:

1. **Molarity = Number of Moles of Solute/Liter of Solution**

MOLALITY

- is the number of moles of solute dissolved in 1kg of solvent.
- UNIT: mole/kg or m

FORMULA:

1. **Molality = Moles of Solute/Kilograms of Solvent**

MOLE FRACTION

- Is a way of describing solution composition. It is the ratio of the number of moles of one component of a mixture to the total number of moles of all components.
- Is the ratio of the number of moles of one component (nA) to the total number of moles in the solution (nA + nB).

EXAMPLE PROBLEM (for steps in getting the mole fraction):

What is the mole fraction of the solute in a 40% by mass ethanol ($C2H6O$) solution in water?

$$\text{MM ethanol} = 46 \text{ g/mole}$$
$$\text{MM water} = 18 \text{ g/mole}$$

Step 1: In converting concentration units based on the mass or moles of a solute and solvent or mass percentage, it is useful to assume a certain total mass of solution. Assume there is exactly 100 grams of solution. Because the solution is 40% ethanol ($C2H6O$), it contains 40 grams of ethanol and 60 grams of water.

Step 2: Change the masses of the components ethanol and water to the number of moles.

$$\text{mole ethanol} = \frac{40 \text{ g}}{46 \text{ g/mol}} = 0.87 \text{ mol}$$

$$\text{mole water} = \frac{60 \text{ g}}{18 \text{ g/mol}} = 3.33 \text{ mol}$$

Step 3: Substitute the values obtained in the formula and solve for the mole fraction of the solute ethanol, and the solvent water.

$$\text{mole fraction}_{\text{ethanol}} = \frac{\text{mole}_{\text{ethanol}}}{\text{mole}_{\text{ethanol}} + \text{mole}_{\text{water}}}$$

$$= \frac{0.87 \text{ mol}}{0.87 \text{ mol} + 3.33 \text{ mol}}$$

$$= 0.207 \text{ mol}$$

LESSON 6: COLLIGATIVE PROPERTIES OF NONELECTROLYTES LESSAND ELECTROLYTE SOLUTIONS

Vocabulary

Coligare

- Tied together

Electrolyte

- Is a compound that conducts an electric current when it is an aqueous solution or melted.

Nonelectrolyte

- Is a compound that does not conduct an electric current in either aqueous solution or in the

COLLIGATIVE PROPERTY

- are properties of a solution that depend only on the number and not on the identity of the solute particles.
  a. Vapor pressure lowering
  b. Boiling point elevation
  c. Freezing point depression

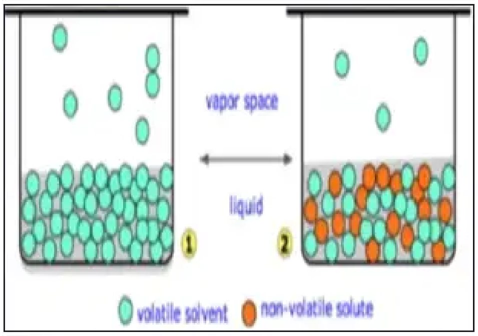
EFFECT OF SOLUTE CONCENTRATION ON THE COLLIGATIVE PROPERTIES OF SOLUTIONS

The concentration or amount of nonvolatile solute in the solution has an effect on the colligative properties of solutions. The effect would depend on the ratio of the number of particles of solute and solvent in the solution and not on the identity of the solute.

EFFECT OF ELECTROLYTE AND NONELECTROLYTE AND NONELECTROLYTE ON COLLIGATIVE PROPERTIES SOLUTIONS

a. **Vapor Pressure Lowering**

- Vapor pressure is a direct measure of escaping tendency of molecules. A pure liquid (solvent) in a closed container will establish equilibrium with its vapor. And when that equilibrium is reached, the pressure exerted by the vapor is called the vapor pressure.



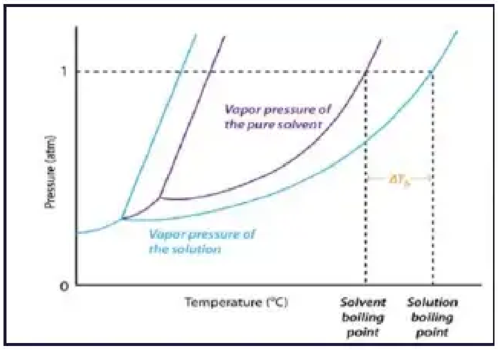
b. **Boiling Point Elevation**

- The addition of a nonvolatile solute lowers the vapor pressure of the solution; consequently the temperature must be raised to restore the vapor pressure of the solution to the value conforming to the pure solvent. Specifically, the temperature at which the vapor pressure is 1 atm will be higher than the normal boiling point by an amount known as the **boiling point elevation.**

FORMULA:

1. **$\Delta T_b = i \times K_b \times m$**

   Where,
   i is the Van't Hoff factor,
   Kb is the ebullioscopic constant,
   m is the molality of the solute.

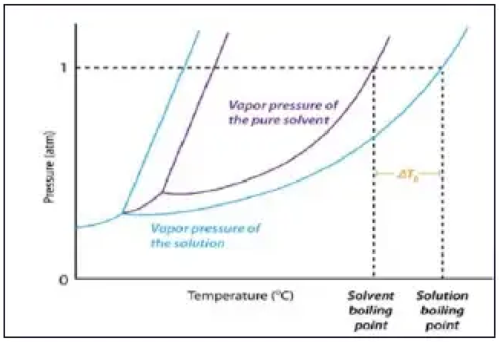


ADDITIONAL INFORMATION:

Van't Hoff factor

- can be defined as the ratio of the concentration of particles formed when a substance is dissolved to the concentration of the substance by mass.

THE LOWERING OF THE VAPOR PRESSURE IN A SOLUTION CAUSES THE BOILING POINT OF THE SOLUTION TO BE HIGHER THAN PURE SOLVENT



c. **Freezing Point Depression**

- The freezing point of a substance is the temperature at which the solid and liquid forms can coexist indefinitely, at equilibrium. Under these conditions molecules pass between the 2 phases at equal rates because their escaping tendencies from the two phases are identical.
- At a given temperature, if a substance is added to a solvent like water, the solute-solvent interactions prevent the solvent from going into the solid phase, requiring the temperature to decrease further before the solution will solidify. Meaning, more energy must be removed from the solution in order to freeze it and the freezing point of the solution is power than that of the pure solvent.

FORMULA:

1. **$\Delta Tf = Kf \times m$**
   Where,
   Kf is the molal freezing point depression constant., a constant that is equal to the change in the freezing-point for a 1 molal solution of a nonvolatile molecular solute,
   Tf is the freezing point depression,
   m is the molality of the solute

ADDITIONAL INFORMATION:

Which would increase the boiling point of water more: salt or sugar? Why?

- More heat energy was needed to break the ionic bond in the salt solution than the covalent bond in the sugar solution so the boiling point of **salt solution** was higher.

When a solute is added to a solvent forming a solution, what happens to the colligative properties of the liquid?

- Both change in boiling and freezing point may occur if a nonvolatile solute is added to the solvent. When solute is added to a solvent like water, its boiling point increases. However, the effect is opposite as regards freezing point.

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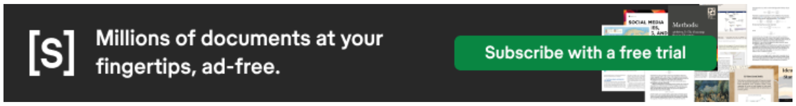

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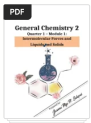
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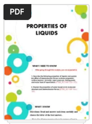
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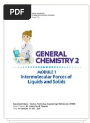
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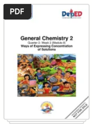

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