

CHEMISTRY 111 LECTURE

EXAM II Material

Part 1 LIQUIDS, SOLIDS, AND INTERMOLECULAR FORCES

Chapter 11

I. Solids, Liquids, and Gases: A Molecular Comparison

PROPERTIES OF GASES

- A. Gas particles are far apart from each other - there is no attraction between particles
- B. Gases have an indefinite shape.
- C. Gases have a low density
- D. Gases are very compressible
- E. Gases exert pressure equally in all directions on the walls of a container.
- F. Gases have a high flow rate. Gases mix spontaneously and completely with one or more other gases.

PROPERTIES OF SOLIDS

- A. A Solid contain particles which are very close to each other-there are very large attractive forces between particles
- B. Solids have a definite shape. Solids maintain its shape regardless of the container they are in.
- C. Solids in general have a high density
- D. Solids are not compressible (ornegligibility)
- E. Solids do not flow

PROPERTIES OF LIQUIDS

- A. Liquids contain particles which are (somewhat) close to each other - there are attractive forces between particles
- B. Liquids have a definite shape. Liquids maintain the shape of the bottom of the container.
- C. Liquids in general have a medium density
- D. Liquids are not compressible (ornegligibility)
- E. Liquids have a medium flow rate

II. Intra and Intermolecular (particle) forces-Review

A.

B. Intramolecular (particle) forces

C. Intermolecular (particle) forces

The attractive forces between particles.

Types

1. Dispersion Forces (London forces) The attraction between atoms and nonpolar molecules. London forces are very weak electrostatic forces of attraction between molecules with "temporary" dipoles.

2. Dipole-Dipole interaction:
Dipole - dipole interactions are electrostatic attractions between polar molecules

3. Hydrogen bonds:
A hydrogen bond is a relatively strong dipole-dipole attractive force between a hydrogen atom and a pair of nonbonding electrons on a F, O, or N atom

3. Ion-Dipole bonds:
A Ion-Dipole bonds occur when an ionic compound is mixed with a polar compound. – such as : aqueous solution!!

II. Electronegativity and bonding- Review (if needed)

A. Electronegativity - The measure of the attractive force that an atom has for its shared electrons. In general, electronegativity increases left to right and bottom to top on the periodic table.

B. Electronegativities and bond polarity

1. Covalent Bonds

a. Non polar covalent bonds - differences in the electronegativities is ≤ 0.4

b. Polar covalent bonds - differences in the electronegativities is between 0.5 -1.7

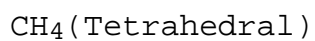
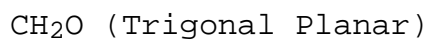
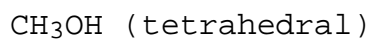
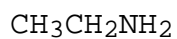
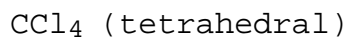
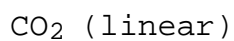
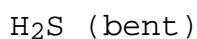
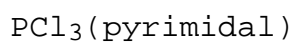
2. Ionic bonds

Differences in electronegativities is > 1.7

Exercises

H with N, O, or H bond:	H - bond / Dipole-Dipole/ Dispersion
Polar molecules:	Dipole- Dipole/ Dispersion
Non polar molecules or Atoms:	Dispersion & Strength of IMF \propto molar mass

1. What is the major type of intermolecular forces for the following



2. Which has the strongest intermolecular forces?

I_2 or Br_2 or Cl_2

CH_3CH_3 or $CH_3CH_2CH_2CH_2CH_2CH_2CH_3$

III Intermolecular Forces in Action Surface Tension, Viscosity, and Capillary Action

A. SURFACE TENSION

1. Definition: The Intermolecular forces at the surface of a liquid between molecules form a tension.
- 2.

B. VISCOSITY

1. Definition: Resistance of a liquid to flow

B. EVAPORATION OF LIQUIDS

1. Definition: Evaporation is the change from the liquid state to the gaseous or vapor state.
Liquid + heat \rightleftharpoons Vapor

Evaporation of Liquids. cont'd

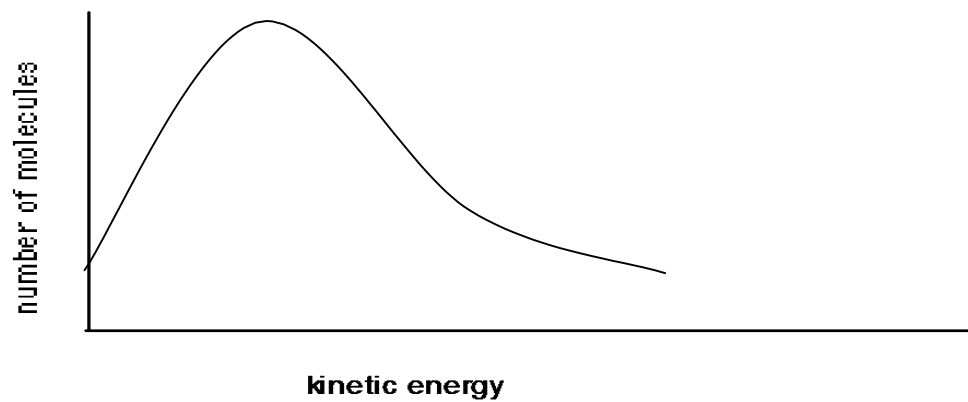
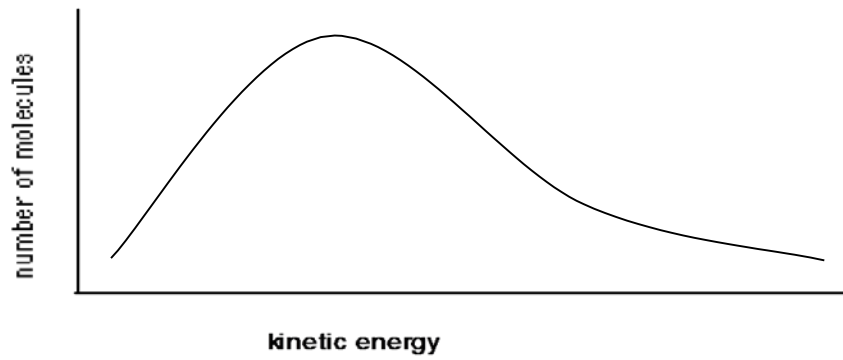
2. Effect of intermolecular forces on evaporation

Examples:

3. Effect of temperature on evaporation

4. Effect of surface area on the rate of evaporation

5. Effect of evaporation on the average kinetic energy



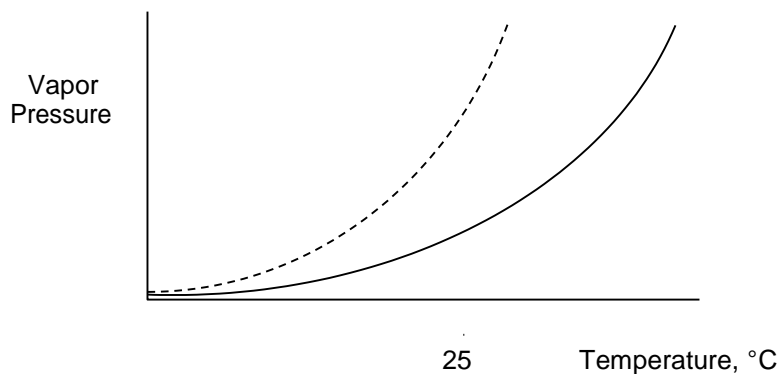
C. VAPOR PRESSURE OF LIQUIDS

The vapor pressure of a liquid is the pressure exerted by the vapor above a liquid.

Dynamic equilibrium: rate of evaporation = rate of condensation

1. Effect of intermolecular forces on vapor pressure

2. Effect of temperature on vapor pressure



*Note: The vapor pressure of a liquid is independent of the volume of the container, provided that there is some liquid present so that equilibrium is established

D. BOILING POINT

The boiling point of a liquid is the temperature at which a liquid is changed to a gas within the liquid (bubbles formed underneath the surface). The **boiling point** is the temperature at which the vapor pressure equals the atmospheric pressure (external pressure).

1. The **normal boiling point** is the temperature at which the vapor pressure equals the atmospheric pressure of 1 atm.

2. Effect of intermolecular forces on the boiling point

IV Changes of State-Review

A. HEAT CAPACITY [Specific heat]

The amount of heat required to raise the temperature of 1 g of a substance exactly 1°C.

Example: How many degrees Celsius will the temperature rise if 25 g ether absorbs 160. cal of energy.

$$\text{Specific heat ether} = \frac{0.529 \text{ cal}}{\text{g } ^\circ\text{C}}$$

B. SUBLIMATION AND FUSION - ENERGY AND CHANGE OF STATE

Energy (as heat) is either lost or absorbed when a substance changes its state

Solid \longrightarrow Liquid

Liquid \longrightarrow Gas

Solid \longrightarrow Gas

C. HEAT OF VAPORIZATION- The quantity of heat needed to convert a liquid at its boiling point to the gaseous state.

$$\text{Ex. } \Delta H_{\text{vap}} = \frac{2.26 \text{ KJ}}{\text{g}}$$

D. HEAT OF FUSION - The quantity of heat needed to convert a solid at its melting point to the liquid state.

$$\text{Ex. } \Delta H_{\text{fus}} = \frac{3.35\text{J}}{\text{g}}$$

IV HEATING CURVE FOR WATER

Problem. How much energy [Heat in kilojoules] is needed to convert 500.0 g of ice at -15.0°C to steam at 105.0°C ?

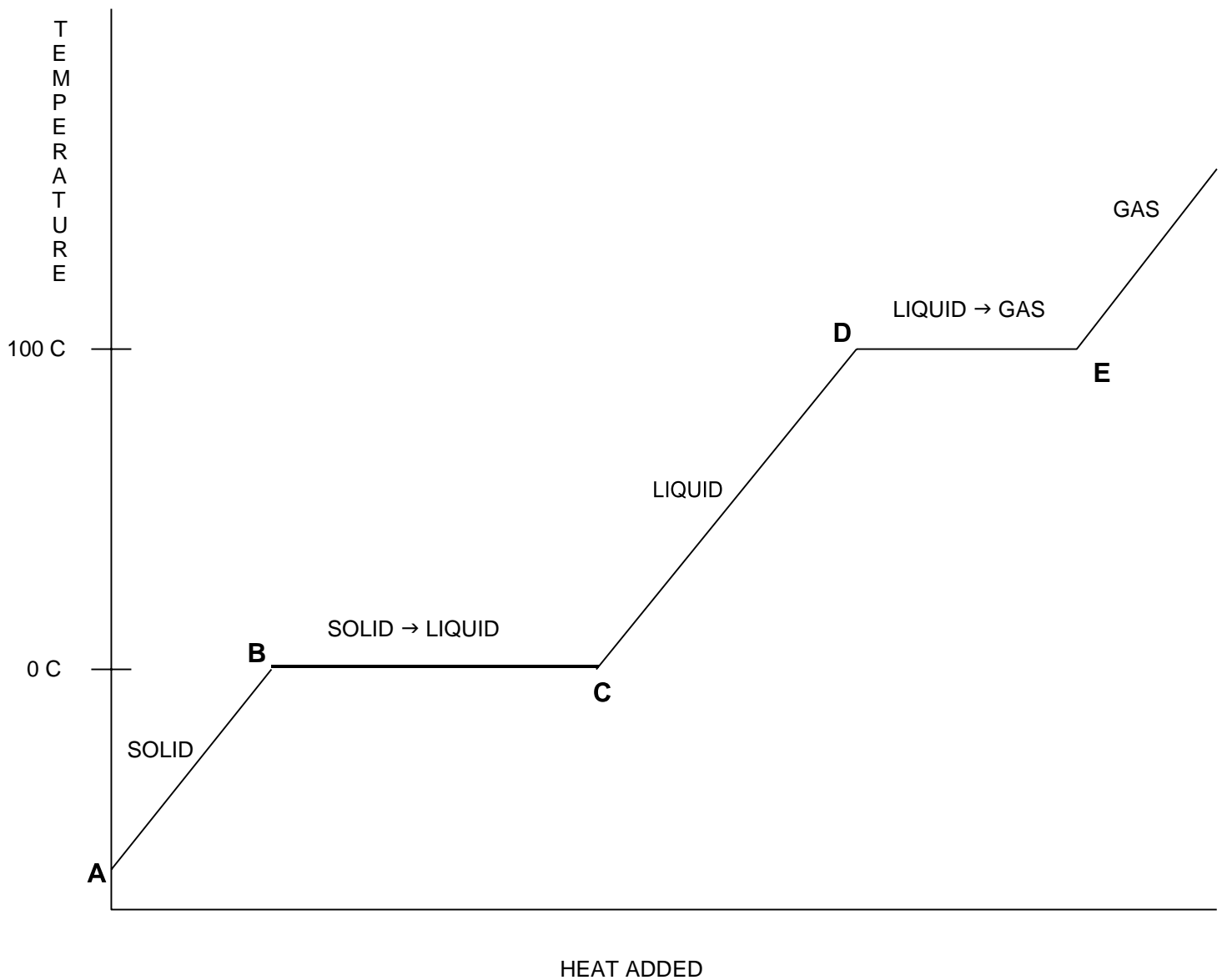
$$H_{\text{fusion}} = \frac{335 \text{ J}}{\text{g}} \quad H_{\text{vap}} = \frac{2.26 \text{ kJ}}{\text{g}}$$

$$\text{Specific heat of ice} = \frac{2.10 \text{ J}}{\text{g } ^{\circ}\text{C}}$$

$$\text{Specific heat of water} = \frac{4.18 \text{ J}}{\text{g } ^{\circ}\text{C}}$$

$$\text{Specific heat of steam} = \frac{2.0 \text{ J}}{\text{g } ^{\circ}\text{C}}$$

GRAPH:



Calculations:

A→B HEATING A SOLID

B→C SOLID → LIQUID

C→D HEATING A LIQUID

D→E LIQUID → GAS

E→F HEATING A GAS

TOTAL HEAT ADDED

E. CRITICAL TEMPERATURE, CRITICAL PRESSURE, AND CRITICAL POINT

To convert a gas to a liquid, you must:

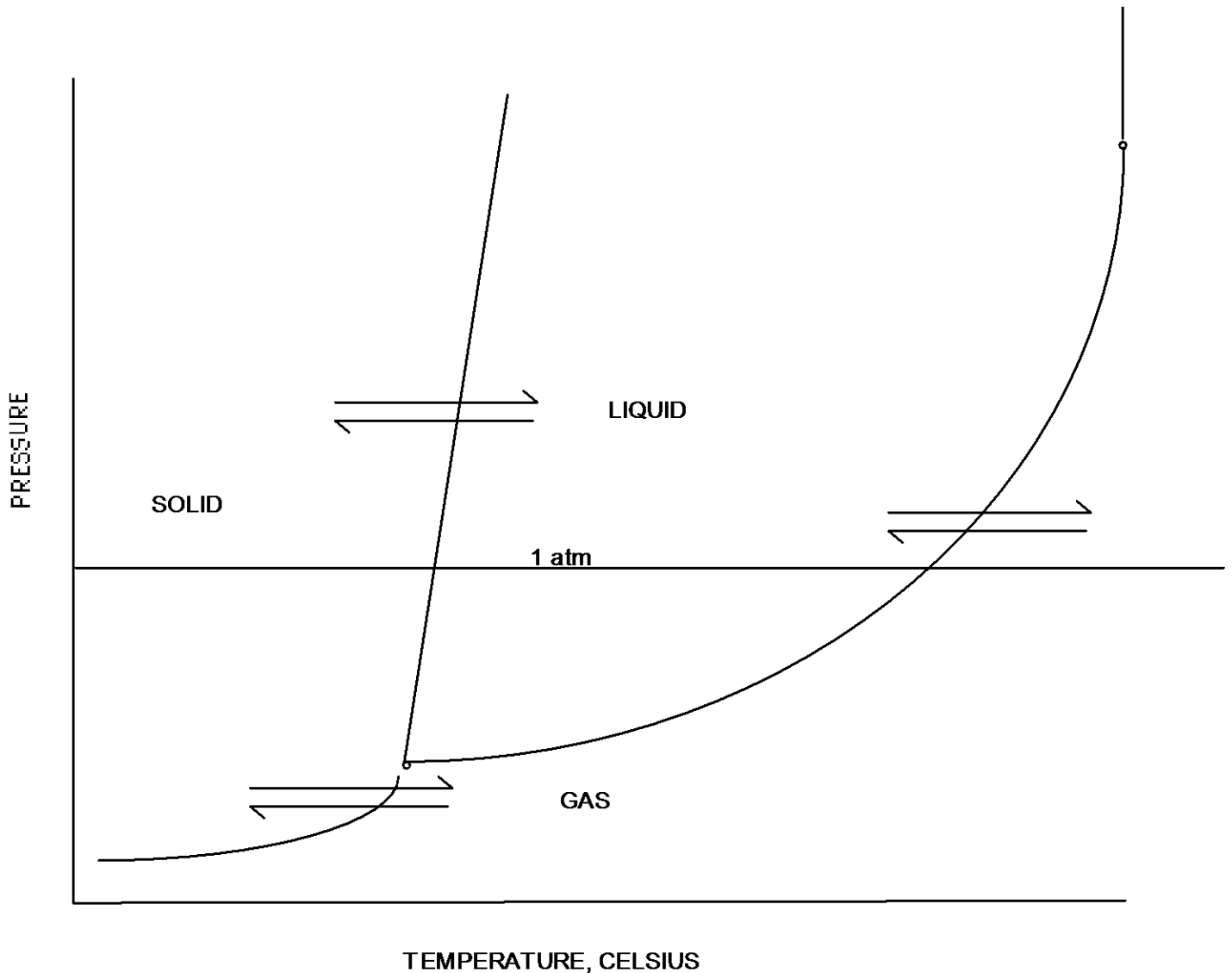
- a. Lower the temperature
and/ or
- b. Increase the pressure

1. The **Critical temperature** of a gas is the highest temperature above which a substance can only exist as a gas. Above its critical temperature, no increase in pressure will convert the gas into a liquid.

2. The **Critical pressure** is the minimum vapor pressure at the critical temperature. It is the minimum pressure required to convert the gas into a liquid at its critical temperature.

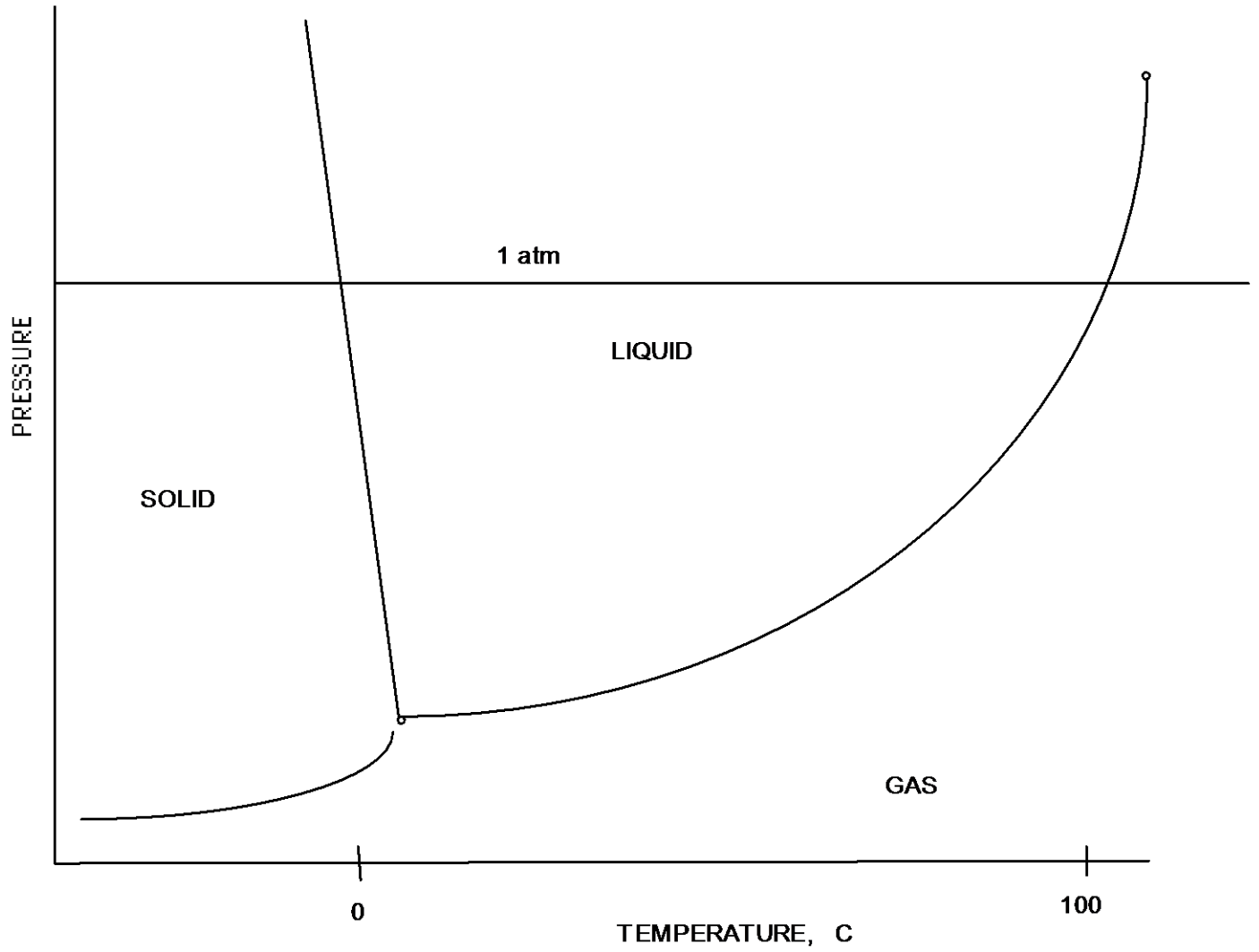
V. PHASE DIAGRAMS

A phase diagram is a graph of pressure versus temperature that shows the region of stability for each of the physical states. The three line segments show the combinations of pressure and temperature at which any two phases exist in equilibrium.

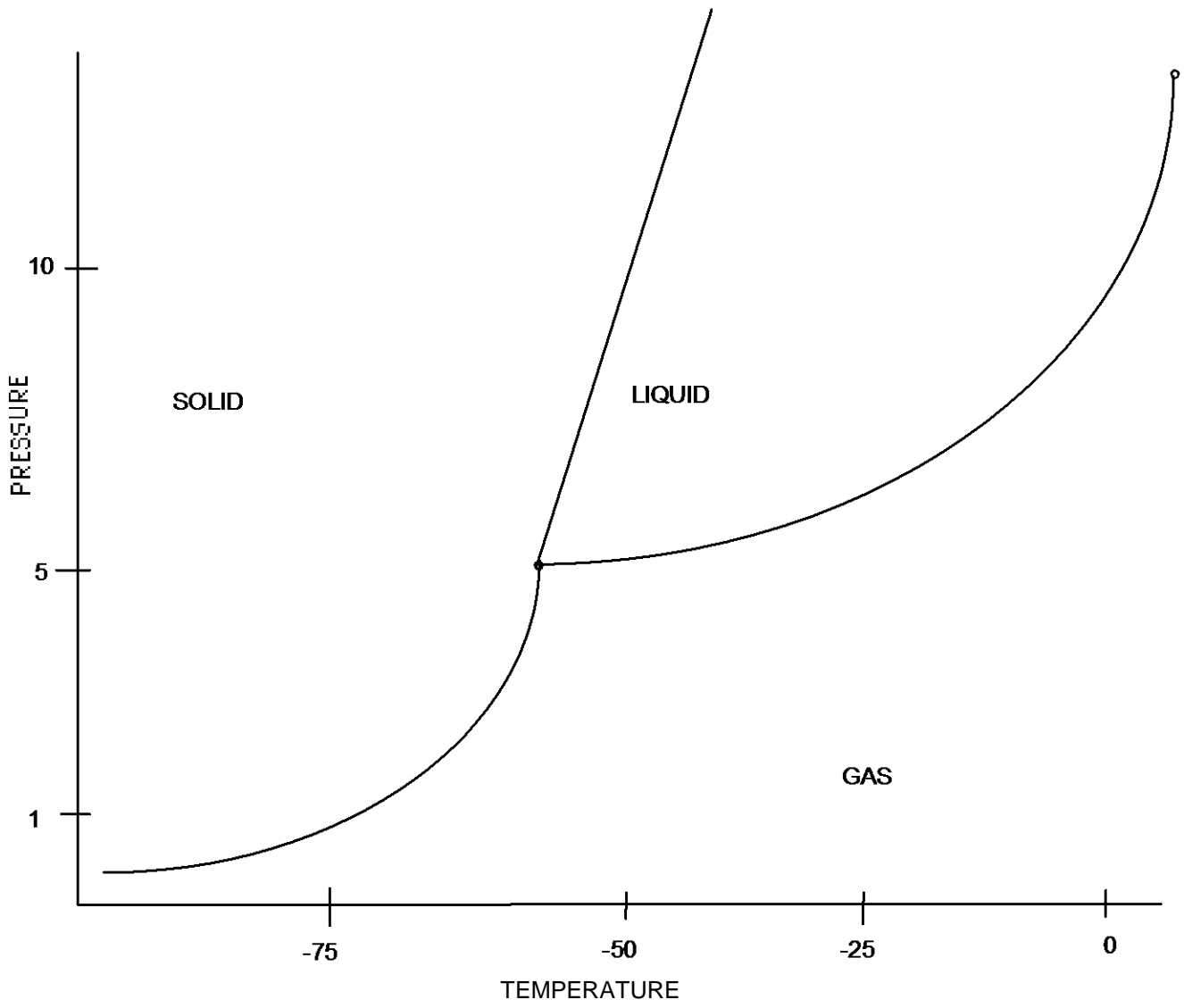


1. The line that separates the gas from the liquid is the vapor pressure curve
2. The **Triple point** is the pressure and temperature at which solid, liquid, and gas all coexist at the same time.
3. The liquid-gas equilibrium line ends at the critical point. Above the critical temperature only a gas exists.

PHASE DIAGRAM FOR H₂O



PHASE DIAGRAM FOR CO₂



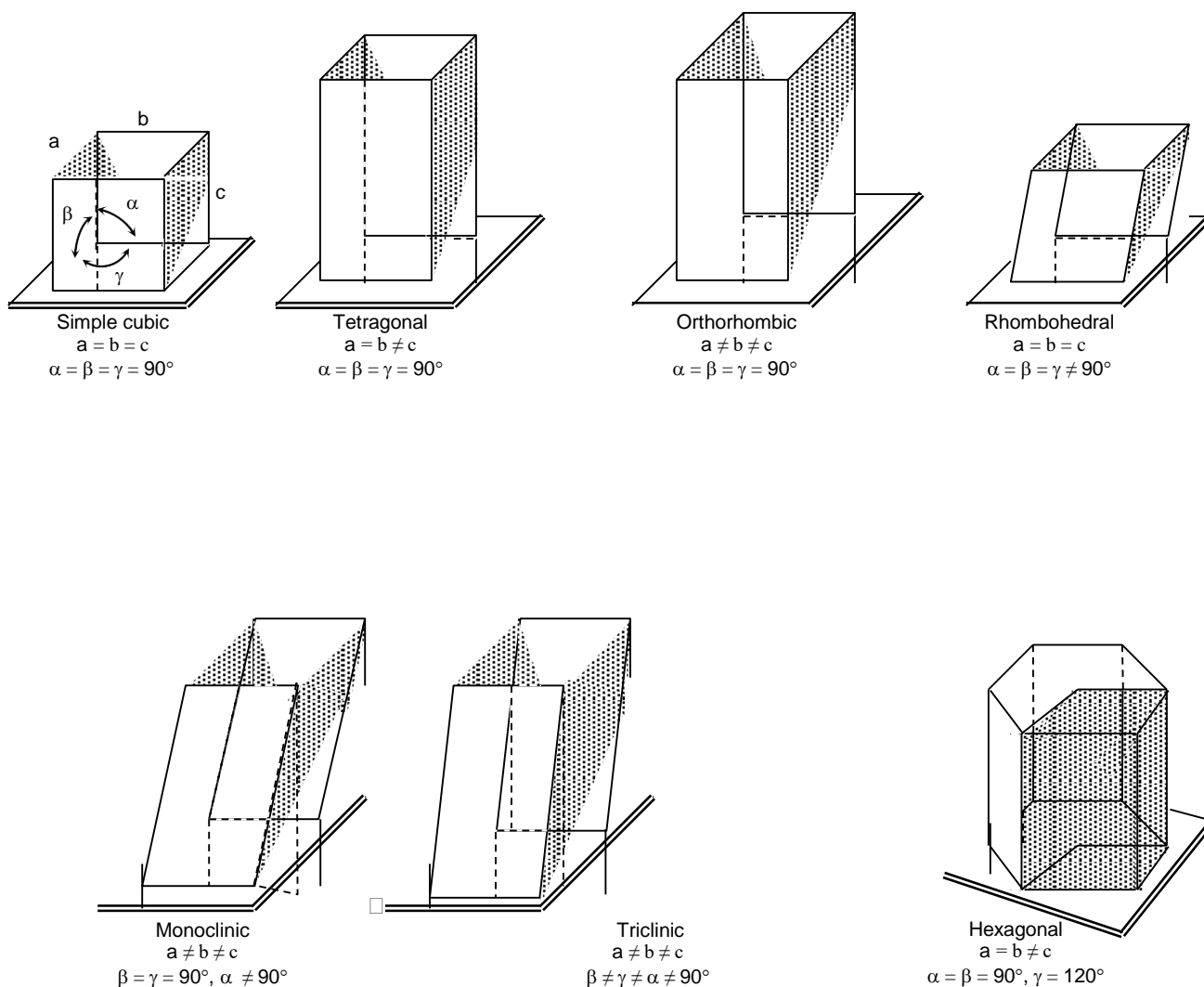
VI. Crystalline Solids: Unit Cells and Basic Structures

A. Types

1. Crystalline Solid
Crystalline solids have a highly ordered arrangement of particles (ions, atoms, and molecules)
2. Amorphous Solid
Amorphous solids have considerable disorder in their structure.

B. Crystalline Solids

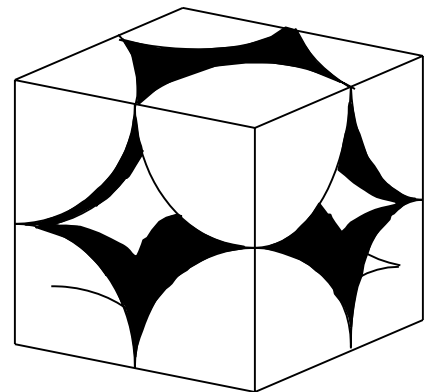
1. Crystalline lattice
A three dimensional array of lattice points in a pattern that defines a crystal.
2. Unit Cell
The Unit cell is the basic repeating unit of the lattice.
3. Coordination Number
The coordination number of a particle in a crystal is the number of nearest neighbors.
4. Lattice Points
The points in a lattice occupied by atoms, ions or molecules.
5. Kinds of Lattice



5. Geometry

6. Simple Cubic

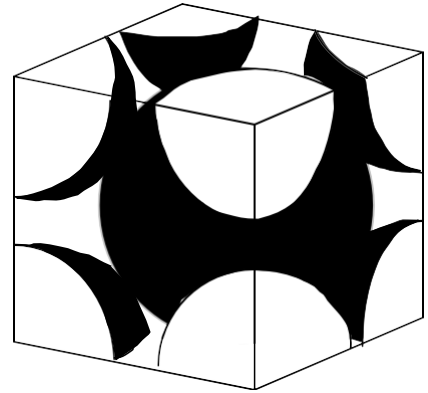
a. Calculate the volume of the unit cell



b. Calculate the volume of the spheres

c. Calculate the percent occupied (packing efficiency)

Body Centered Cubic

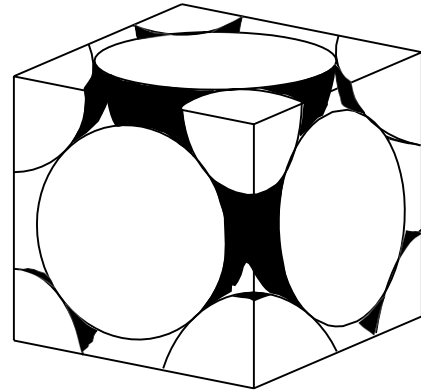


a. Calculate the volume of the unit cell

b. Calculate the volume of the spheres

c. Calculate the percent occupied (packing efficiency)

6. Face Centered Cubic



a. Calculate the volume of the unit cell

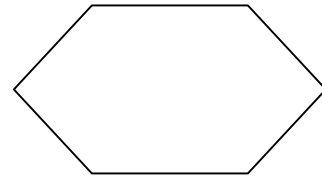
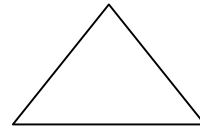
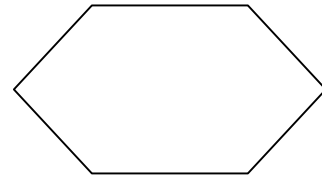
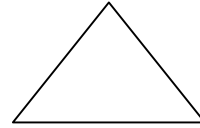
b. Calculate the volume of the spheres

c. Calculate the percent occupied (packing efficiency)

8. **Packing Efficiency**

a. Closest cubic packing

b. Hexagonal



Part V Types of Solids

	<u>IONIC SOLIDS</u>	<u>MOLECULAR SOLIDS</u>	<u>NONBONDING SOLIDS</u>	<u>NETWORK COVALENT</u>	<u>METALLIC SOLIDS</u>
PARTICLES AT LATTICE POINTS	cations and anions [positive and negative ions]	polar or nonpolar covalent molecules	atoms	atoms	metallic cations [positive metallic ions]
BONDING BETWEEN PARTICLES (attractive forces)	electrostatic attractions [ionic bond]	intermolecular forces dipole-dipole, hydrogen bond, london forces	dispersion forces	covalent bonds	metallic bonds between positive metal ions and mobile electrons
PROPERTIES	high M.P. high B.P. nonconductors brittle hard/brittle	moderate to low M.P. moderate to low B.P. nonconductors or semi conductors soft	low M.P. low B.P. nonconductors	high M.P. high B.P. nonconductors brittle hard	variable M.P. variable B.P. malleable and ductile
EXAMPLES	NaCl KBr LiNO ₃	H ₂ O SO ₂ NH ₃ C ₂ H ₅ OH	Xe Ar	C (diamond) SiC SiO ₂	Au Cu Fe Na Hg

I. Solutions are homogeneous mixtures

II. Components of a solution:

III. Solubility-Review:

The amount of solute that dissolves in a given amount of solvent at a given T⁰ and Pressure

a. In: $\frac{\text{g solute}}{100 \text{ g solvent}}$

ex. $\frac{79.5 \text{ g NaBr}}{100 \text{ g H}_2\text{O}}$ vs $\frac{0.00015 \text{ g Fe(OH)}_3}{100 \text{ g H}_2\text{O}}$

b. Past solubility ⇒ Additional solute will not dissolve

c. Concentration of solutions

- 1) Dilute solutions contains a small amount of solute
- 2) Concentrated solutions contains a large amount of solute

d. Solubility terms for **solids** as the solute

- 1) In an **unsaturated solution** additional solute will dissolve
- 2) In a **saturated solution** additional solute will **not** dissolve
- 3) A **supersaturated solution** is a solution that has been prepared to hold more solute than its solubility limit

e. Solubility terms for solids as the solute

Saturated, Unsaturated, or supersaturated solution?

f. Solubility terms for **liquids** as the solute

- 1) Miscible-2 liquids that form a solution in all proportions
- 2) Immiscible - 2 liquids that do not form a solution

SOLUBILITY RULES FOR IONIC COMPOUNDS-review

<u>Ion contained in the Compound</u>	<u>Solubility</u>	<u>Exceptions</u>
Group IA	soluble	
NH ₄ ⁺	soluble	
C ₂ H ₃ O ₂ ⁻	soluble	
NO ₃ ⁻	soluble	
Cl ⁻ , Br ⁻ , and I ⁻	soluble	Ag ⁺ , Pb ²⁺ , Hg ₂ ²⁺
SO ₄ ²⁻	soluble	Ca ²⁺ , Sr ²⁺ , Ba ²⁺ , Pb ²⁺

CO ₃ ²⁻ , PO ₄ ³⁻ , SO ₃ ²⁻	insoluble	group IA and NH ₄ ⁺
S ²⁻	insoluble	group IA, IIA, and NH ₄ ⁺
OH ⁻	insoluble	group IA, Ca ²⁺ , Ba ²⁺ , Sr ²⁺

STRONG ACIDS AND BASES

STRONG BASES

LiOH	NaOH
KOH	RbOH
CsOH	Sr(OH) ₂
Ba(OH) ₂	Ca(OH) ₂

STRONG ACIDS

HNO ₃	H ₂ SO ₄
HClO ₄	HCl
HBr	HI

f. Factors that Effect Solubility-review

1) Temperature -

a. Solids - In general the solubility of a solid increases with increasing temperature (depends if the solution process is an endo- or exothermic process)

b. Gases - In general the solubility of a gas decreases with increasing temperature.

2) Pressure

a. solids and liquids - Pressure has little / no effect on the solubility of solids.

b. Gases - The solubility of a gas increases when the partial pressure of the gas over the solution is increased (Henry's law)

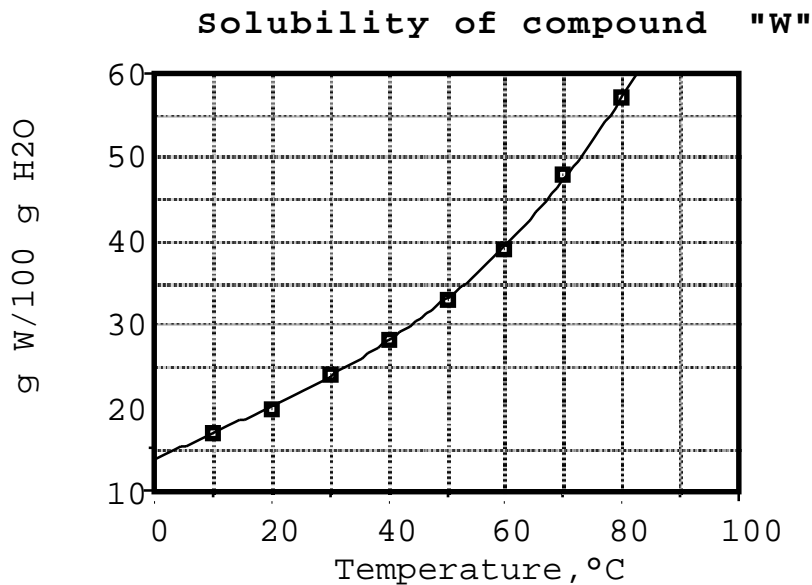
3) Nature of the Solute/Solvent "like dissolves like"

a.

b.

c.

g. Solubility curves-review



IV. Solution Formation

The solution process:

The attraction between ions and water is called **hydration**. This is due to an ion-dipole interaction where the polar water molecules are attracted to the surface of an ionic crystalline lattice. The ion- dipole attractions formed allow the ions to escape from the crystal and are essentially "dissolved" in water. All ions are hydrated in water solution.

V. Heat of solution

Heat can be released or absorbed when a solute dissolves in a solvent (the solvation process).

The enthalpy change, ΔH_{soln} , depends upon the energy required to break solute-solute bonds and the energy released when solute-solvent bonds are formed.

A. An exothermic process:

B. An endothermic process:

VI. Concentrations-Review

A. Percent solute

1. % by weight = $\frac{\text{g solute}}{\text{g solution}} \times 100$

$$\frac{\text{g solute}}{100 \text{ g solution}}$$

2. % by volume = $\frac{\text{ml solute}}{\text{g solution}} \times 100$

$$\frac{\text{ml solute}}{100 \text{ ml solution}}$$

3. Wt-Vol % = $\frac{\text{g solute}}{\text{ml solution}} \times 100$

$$\frac{\text{g solute}}{100 \text{ ml solution}}$$

B. Molarity = $M = \frac{\text{moles solute}}{\text{Liter solution}}$

Problem: What is the molar concentration of a solution that has 10.3 g of sodium bromide in 251 mL of solution?

C. Molality = $\frac{\text{moles solute}}{\text{Kg solvent}}$

Problem: What is the molality of a solution that has 10.3 g sodium bromide that has been dissolved in 300. mL of water?

D. Mole Fraction = $X_A = \frac{\text{moles } A}{\text{total moles of solution}}$

Problem: What is the mole fraction of sodium bromide when the solution contains 10.3 g sodium bromide dissolved in 300. mL of water?

VII Dilutions:

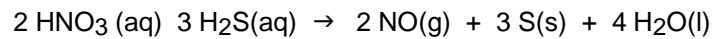
$$M_1V_1 = M_2V_2$$

Dilution Problem

1. 25ml of a 8.0 M HCl solution is diluted to 1 liter. What is the final molarity?

Solution Problems-Review

1. How many grams of S can be produced from the reaction of 30.0 mL of 12.0 M HNO₃ with an excess of 0.035 M H₂S.



2. What is the molarity of a 50.0 mL sample of sulfuric acid that will completely react with 40.0 mL of 0.200 M Mg(OH)₂?

3. The density of a 88.00 % (wt/wt) methanol (CH₃OH) solution is 0.8274 g/ml. What is the molarity of the solution?

4. An aqueous solution of acetic acid is 0.796 m and has a density of 1.004 g/ml. What is the molarity of the solution?

5. Calculate the molality of a 14.0 % by mass nitric acid solution.

Part 6 ELECTROLYTES-Review

I. **ELECTROLYTES**-A substance that is a conductor of electricity in water due to the movement of ions in solution:

A. Strong, Weak, and Nonelectrolytes

1. Strong Electrolytes:
solution.

- a. Substances which are strong conductors of electricity in an aqueous solution.
- b. Substances that are 100% ionized in solution
- c. Substances which are strong electrolytes:
 - (1) Soluble ionic compounds
 - (2) Strong Acids
 - (3) Strong Bases

2. Weak Electrolytes:

- a. Substances which are weak conductors of electricity in an aqueous solution.
- b. Substances which ionize very little in solution
- c. Substances which are weak electrolytes:
 - (1) Weak Acids
 - (2) Weak Soluble Bases
 - (3) Slightly soluble ionic compounds

3. Nonelectrolytes:

- a. Substances which do not conduct electricity in solution.
- b. Substances which do not ionize in solution.
- c. Substances which are nonelectrolytes:

Soluble substance that only exists as molecules in water

Part 7 NET-IONIC EQUATIONS-Review

Net Ionic equations shows the species that are reacting in solution

Un-ionized equation - the bookkeeping equation

Total ionic equation - Shows substances in their predominant form

Net-ionic equation - Shows the only species that underwent a chemical reaction. [Spectator ions have been eliminated]

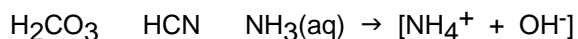
How to write net-ionic equations

1. Write a balanced equation (correct chemical formulas)
2. Write a total ionic equation:
 - a. Write the following in the ionized form:

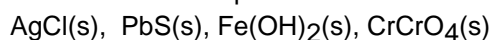
<u>Soluble Salt</u>	FeCl ₂ (aq)	Write As: Fe ²⁺ (aq) + 2Cl ⁻ (aq)
<u>Strong Acid</u>	HCl(aq)	H ⁺ (aq) + Cl ⁻ (aq)
<u>Strong Base</u>	NaOH(aq)	Na ⁺ (aq) + OH ⁻ (aq)

- b. Write the following in the unionized form:

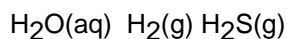
(1) Weak acids and bases:



(2) Insoluble ionic compounds



(3) Molecules



3. Write the net-ionic equation by eliminating all spectator ions. (The unreacting species)
The net-ionic equation must be in the simplest ratio possible

If all species on both sides are spectator ions → N.R.

EXAMPLES:

1. Oxalic acid is poured into a solution of potassium hydroxide.

Molecular equation_____

Total ionic_____

Net ionic_____

2. Solutions of Iron (II) chloride and cesium hydroxide are mixed together

Molecular equation_____

Total ionic_____

Net ionic_____

3. Aqueous sodium phosphate and sulfuric acid are mixed.

Molecular equation_____

Total ionic_____

Net ionic_____

4. . lead (II) cyanide and potassium carbonate solutions are mixed

Molecular equation_____

Total ionic_____

Net ionic_____

Concentration of Ions after a Limiting Reactant Reaction

20.0 mL of a 0.100 M sodium carbonate solution is mixed with 40.00 mL of a 0.800 M nickel (III) chloride solution.

- a) Calculate the number of grams of nickel (III) carbonate produced.
- b) Calculate the molarity of all the species in solution after the reaction has taken place.

The balanced eqn.: $3 \text{Na}_2\text{CO}_3(\text{aq}) + 2 \text{NiCl}_3(\text{aq}) \rightarrow \text{Ni}_2(\text{CO}_3)_3(\text{s}) + 6 \text{NaCl}(\text{aq})$

The total ionic eqn.:

Calculations:

PART B VAPOR PRESSURE OF SOLUTIONS

I. VAPOR PRESSURE

If a solution is composed of more than 1 volatile component, the total pressure of gases about the solution is the sum of the partial pressure of all of the gases:

$$P_{\text{total}} = P_A + P_B + P_C \dots + \dots \text{etc.}$$

Where:

P_A = partial vapor pressure of A

P_B = partial vapor pressure of B

P_C = partial vapor pressure of C

II. PARTIAL PRESSURES - RAOULT'S LAW

The partial pressure of a component is found by:

$$P_A = X_A P_A^\circ$$

Where:

P_A = partial vapor pressure of A over the solution.

P_A° = vapor pressure of A over a pure solution of A

X_A = The mole fraction of A in solution.

THEREFORE:
$$P_{\text{total}} = P_A + P_B + P_C$$

$$P_{\text{total}} = X_A P_A^\circ + X_B P_B^\circ + X_C P_C^\circ$$

III. PROBLEMS

1. VAPOR PRESSURE OF SOLUTION CONTAINING A NONVOLATILE, NONIONIZING SOLUTE

100.0 g of glycerin, $C_3H_8O_3$ (MM= 92.1 g/mol), a nonvolatile nonelectrolyte is added to 200.0 mL of water 25.0°C.

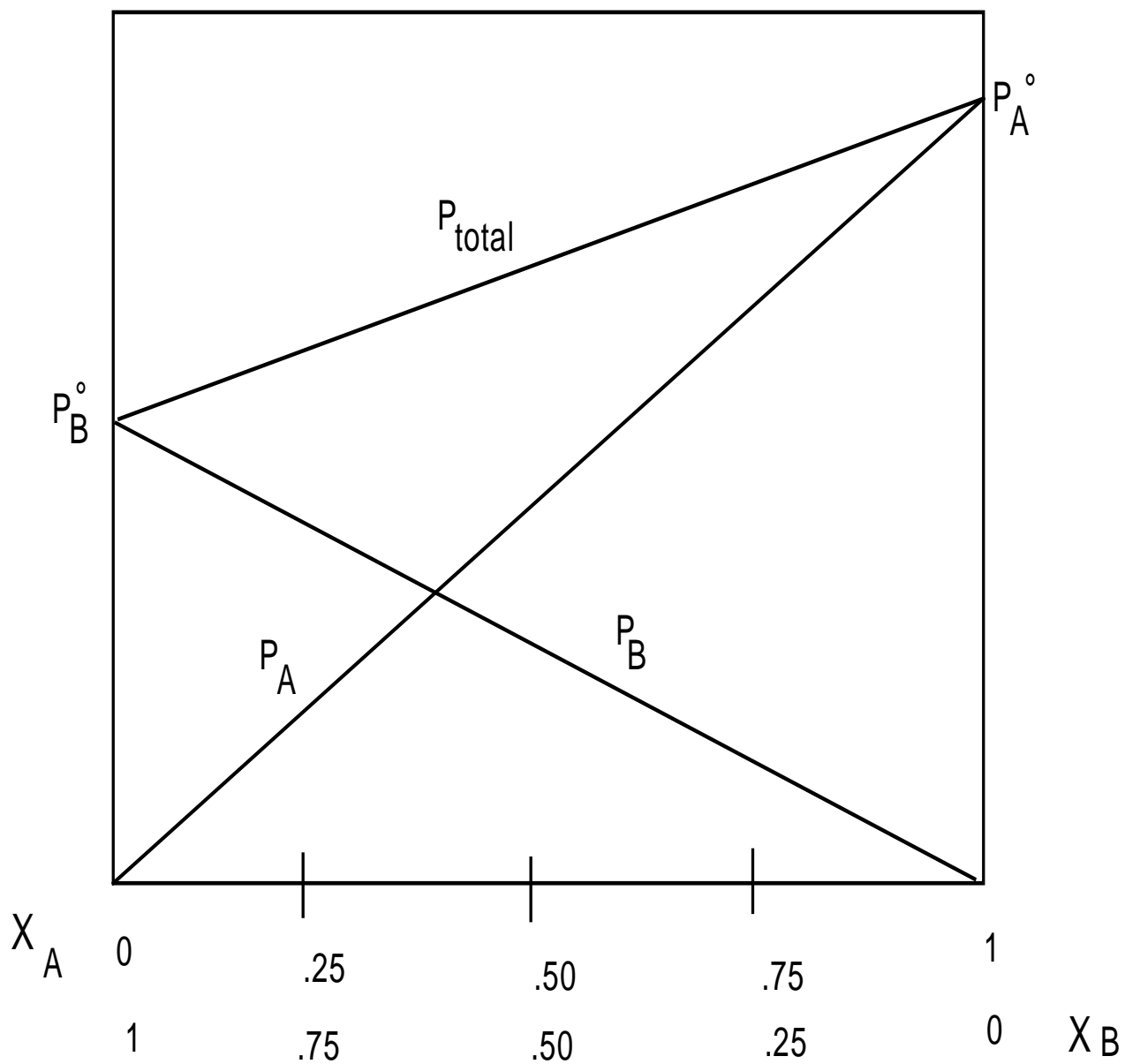
The vapor pressure of pure water at 25.0°C is 23.8 torr. What is the vapor pressure of the resulting solution?

2. VAPOR PRESSURE OF SOLUTION WITH TWO OR MORE VOLATILE COMPONENTS

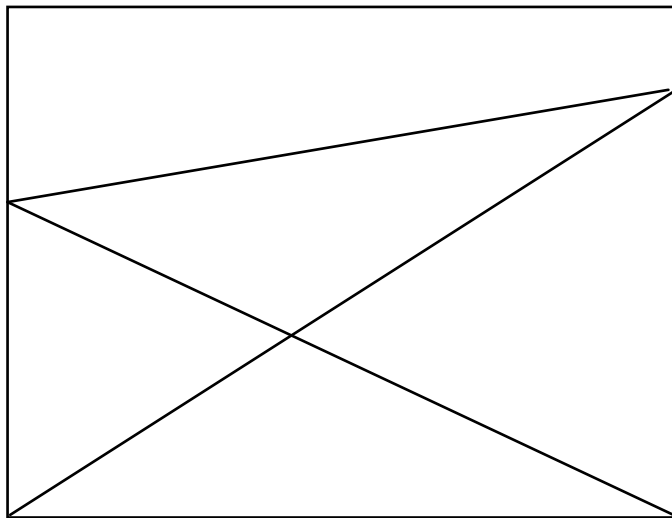
Equal masses (100.0 g) of Ethanol (C_2H_5OH , MM= 46.0) and water (, MM= 18.0) are mixed at 63.5 °C and an ideal solution is formed. At 63.5°C The vapor pressure of ethanol and water is 400.0 torr and 175 torr, respectively. Calculate the Vapor pressure over the solution

IV. RAOULT'S LAW-PARTIAL PRESSURE CURVES

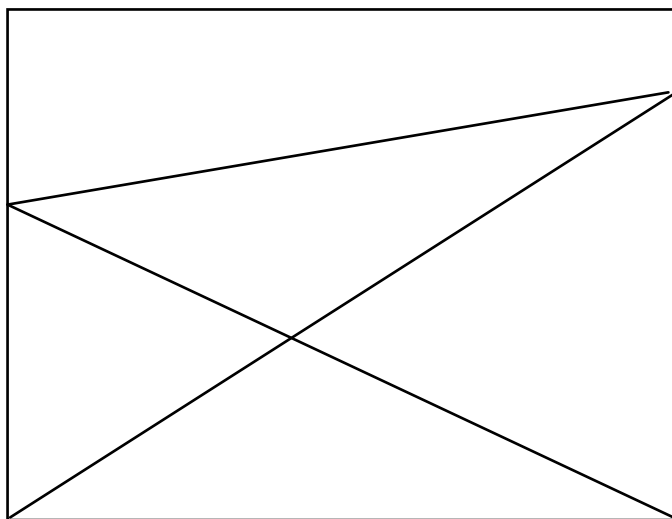
A. An IDEAL SOLUTION follows Raoult's law. The intermolecular forces between component A and B are the same as between A-A and B-B



B. A **POSTIVE DEVIATION FROM RAOULT'S LAW**: The partial pressure of A and B and total pressure above the solution are HIGHER than expected. The intermolecular forces between component A and B are weaker than between A-A and B-B



C. A **NEGATIVE DEVIATION FROM RAOULT'S LAW**: The partial pressure of A and B and total pressure above the solution is LOWER than expected. The intermolecular forces between component A and B are stronger than between A-A and B-B



Part 9 COLLIGATIVE PROPERTIES

The **colligative properties** of solutions are properties that depend primarily upon the number of dissolved particles present (concentrations), rather than upon the nature of these particles.

Colligative properties include: Boiling point elevation, freezing point lowering, and osmotic pressure.

I. BOILING POINT ELEVATION

The boiling point of a liquid is the temperature at which its vapor pressure equals the atmospheric pressure. The normal boiling point is when the vapor pressure equals the atmospheric pressure of 1. The addition of a solute decreases the vapor pressure above the liquid. Therefore, an increase of temperature is required to achieve a vapor pressure of 1 atm. This increase in the boiling temperature when a solute is added is called Boiling Point Elevation.



The boiling point elevation, ΔT_b , is the difference between the boiling point of the solvent and the boiling point of the solution.

$$\Delta T = iK_b m$$

Where: K_b = molal boiling point elevation constant for

the solvent Units of K_f are: $\frac{^{\circ}\text{C kg}}{\text{mole}}$.

m = the molality of the solution = $\frac{\text{moles solute}}{\text{kg solvent}}$

II. FREEZING POINT DEPRESSION

The addition of a solute to a solvent will decrease the freezing point (temperature) of the solvent. The freezing point depression, ΔT , is the difference between the freezing point of the solvent and the freezing point of the solution.

$$\Delta T = iK_f m$$

Where: K_f = molal freezing point depression constant for
the solvent Units of K_f are: $\frac{^{\circ}\text{C kg}}{\text{mole}}$.
 m = the molality of the solution = $\frac{\text{moles solute}}{\text{kg solvent}}$

Problems:

1. An aqueous solution is 0.0222 m glucose. What is the freezing point of this solution?

2. Camphor melts at $179.5\text{ }^{\circ}\text{C}$ with a freezing point depression constant of $40^{\circ}\text{C}/\text{m}$.
- A 1.07 mg sample of an unknown compound was dissolved in 78.1 mg of camphor. The solution melted at 176.0°C . What is the molecular weight of the compound?
 - If the empirical formula of the unknown compound is CH , what is the molecular formula

III. OSMOSIS

Osmosis is the phenomenon of solvent flowing from a high solute concentration through a semi-permeable membrane to a lower solute concentration.

The colligative property, OSMOTIC PRESSURE, is the pressure applied to the solution that stops osmosis from occurring.

$$\pi = iMRT$$

π = osmotic pressure

M = molarity

R = Gas constant

T = temperature

Problem:

Calculate the osmotic pressure at 25°C. When 0.798 g of starch (MM= is 3.24×10^4 on the average) is dissolved to produce 100.0 ml of solution.

IV. COLLIGATIVE PROPERTIES OF IONIC SOLUTIONS

Colligative properties depend only upon the number of particles in the solution. NaCl dissociates in solution into two ions. Therefore the ΔT_f , or ΔT_b , for a 1 M NaCl solution would be twice that of 1 M glucose solution, a nonelectrolyte.

$$\Delta T_b = iK_b m$$

Where: i = the van't Hoff factor

$$\Delta T_f = iK_f m$$

$$\pi = iMRT$$

Type 1 - i is given:

A 0.300 m solution of benzoic acid has an $i = 1.50$. What is the freezing point K_f of water = $1.86 \text{ }^\circ\text{C/m}$

Type 2 - Determine i given the ionic or molecular aqueous compound:

Compound	i

Type 3 - Calculate i .

Part 4 COLLOIDS

A. **COLLOID** is a homogeneous mixture. Particles of one substance are dispersed throughout another substance.

COMPONENTS:

COLLOID

Colloid particle

Continuous phase

Dispersed phase

Dispersing medium

B. TYPES OF COLLOIDS

Colloids are classified according to the state (solid, liquid, gas) of the dispersed phase and the state of the continuous phase.

DISPERSED PHASE	CONTINUOUS PHASE	NAME OF COLLOID	EXAMPLES
LIQUID	GAS	LIQUID AEROSOL	➤ FOG ➤ MIST ➤ CLOUD
SOLID	GAS	SOLID AEROSOL	➤ SMOKE ➤ DUST
GAS	LIQUID	FOAM	➤ BEER FROTH ➤ SAP SUDS ➤ WHIPPED CREAM
LIQUID	LIQUID	EMULSION	➤ MILK (CREAM IN MILK) ➤ MAYONNAISE (OIL DISPERSED IN WATER)
SOLID	LIQUID	SOL	➤ AgCl PPT IN H ₂ O ➤ PAINT ➤ INK ➤ MILK OF MAGNESIA
GAS	SOLID	SOLID FOAM	➤ PUMICE ➤ PLASTIC ➤ MARSHMALLOW
LIQUID	SOLID	SOLID EMULSION	➤ CHEESE ➤ BUTTER ➤ JELLY

COMPARISON: TRUE SOLUTION

COLLOID

SUSPENSION

Particle size: < 1 nm

1 - 100 nm

> 100 nm

C. PROPERTIES OF COLLOIDS

1. TYNDALL EFFECT

The Tyndall effect is the scattering of light by colloidal particles.

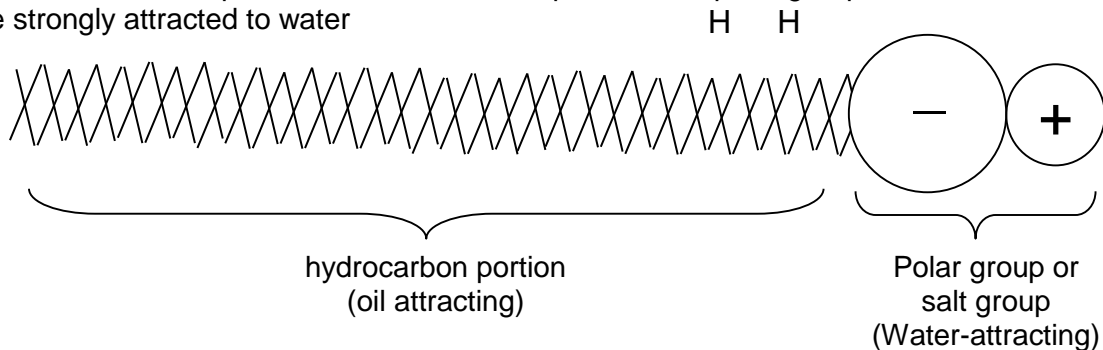
4. RATE OF SETTLING

Very dense dispersed colloidal particles do not sink but may remain dispersed in the dispersing medium indefinitely. Colloidal systems can remain stable indefinitely.

Hydrophilic Colloidal Systems

a. Macromolecules - The colloidal particles are macromolecules which are more attracted to water (the dispersing phase) than each other.

Macromolecules disperse in water due to the presence of polar groups in the molecule which are strongly attracted to water



Macromolecules disperse in water due to the molecules clustering together to form a colloidal particle, where, the polar groups of the molecules are on the outside and are strongly attracted to water (the nonpolar groups are on the inside).

