

States of Matter

■ Intramolecular & Intermolecular Forces

Liquids, Solids, and Phase changes

- An Overview of Physical Changes and Phase Changes
- Quantitative Aspects of Phase Changes
 - Heat Involved in Phase Changes
 - Equilibrium Nature of Phase Changes
 - Phase Diagrams
- Types of Intermolecular Forces
 - Ion-Dipole Forces
 - Dipole-Dipole Forces
 - The Hydrogen Bond
 - Charge-Induced Dipole Forces
 - Dispersion (London) Forces

States of Matter

- Properties of the Liquid State
 - Surface Tension
 - Capillarity
 - Viscosity
- The Uniqueness of Water
 - Solvent Properties
 - Thermal Properties
 - Surface Properties
 - Density of Solid and Liquid Water

States of Matter

- Solid State: Structure, Properties, and Bonding
 - Structural Features of Solids
 - Crystalline Solids
 - Amorphous Solids
 - Bonding in Solids
- Advanced Materials
 - Electronic Materials
 - Liquid Crystals
 - Ceramic Materials
 - Polymers
 - Nanotechnology

States of Matter

- Physical States of Matter – any physically distinct, homogenous part of a system
 - Phases - Types of Physical State
 - Gases
 - Liquids
 - Solids
 - Plasma
- Intermolecular and Intramolecular Forces
 - In all of the 1st 3 phases, Electrostatic Forces among (between) the particles combine with the Kinetic Energy of the particle to create the properties of each phase as well as phase changes
 - Vapor Pressure
 - Boiling Point
 - Melting Point

States of Matter

■ Comparison of gases, liquids, and solids

➤ **Gases**

- Compressible fluids conforming to the shape and volume of container
- Molecules are widely separated

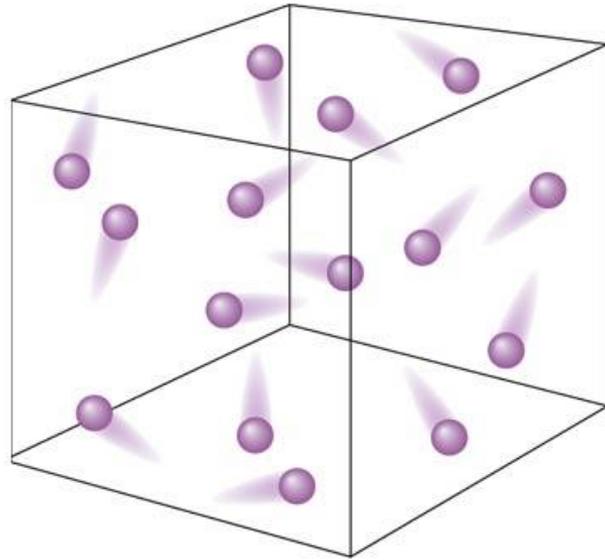
➤ **Liquids**

- Relatively incompressible fluids
- Conform to shape of container, volume limited by surface
- Molecules are more tightly packed

➤ **Solids**

- Nearly incompressible and rigid, maintaining their own shape and volume
Molecules or ions are in close contact and do not move

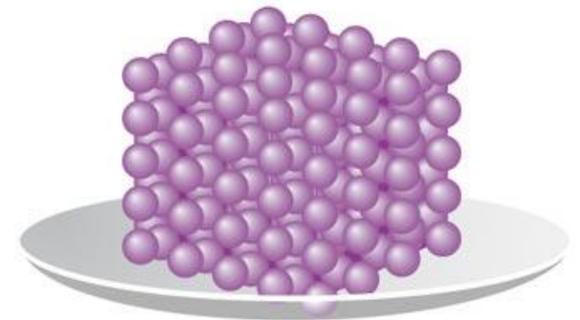
Representation of the States of Matter



Gas



Liquid



Solid

States of Matter

- **Condensation** – Process by which a gas changes to a liquid
- **Vaporization** – Process by which a liquid changes to a gas
 - **Heat of Vaporization (ΔH_{vap})** – Endothermic Enthalpy Changes
- **Freezing** – Process by which a liquid changes to a solid
- **Melting (Fusion)** – Process by which a solid changes to a liquid
 - **Heat of Fusion (ΔH_{fus})** – Endothermic Enthalpy Change
- **Sublimation** – Process by which a solid changes **directly** to a gas
 - Heat of Sublimation (ΔH_{sub})** – Endothermic Enthalpy Change
- **Deposition** – Process by which a gas changes **directly** to a solid

Changes of State

- A change of state or phase transition is a change of a substance from one state to another



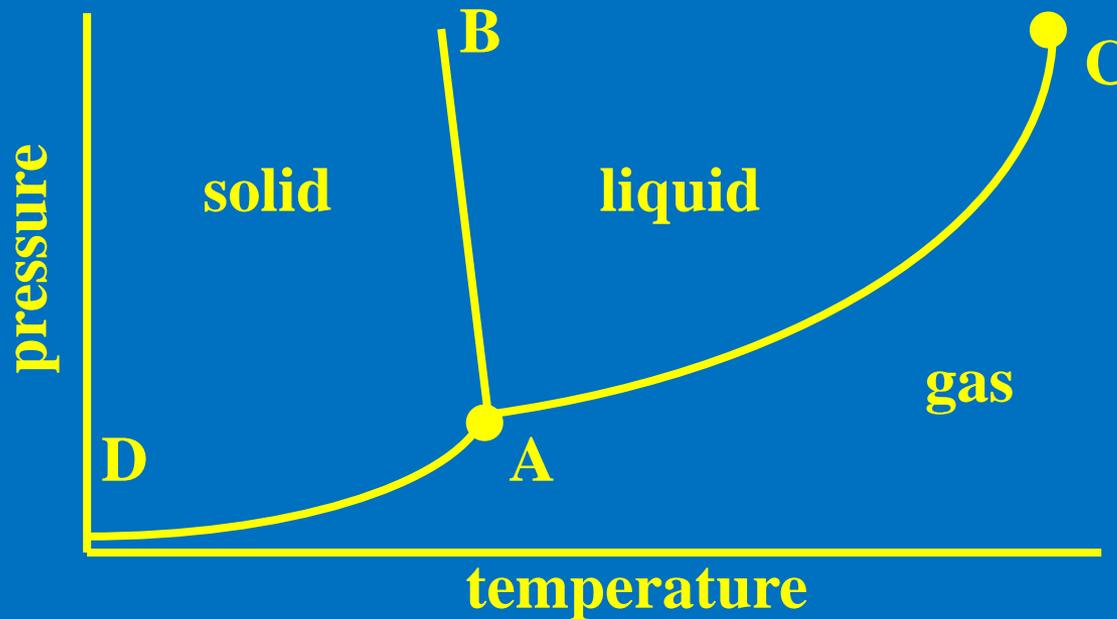
Phase Diagrams

- A **phase diagram** is a graphical way to summarize the conditions under which the different states of a substance are stable
 - The diagram is divided into three areas representing each state of the substance
 - Gas
 - Liquid
 - Solid
 - The curves separating each area represent the boundaries of phase changes

Phase Diagrams

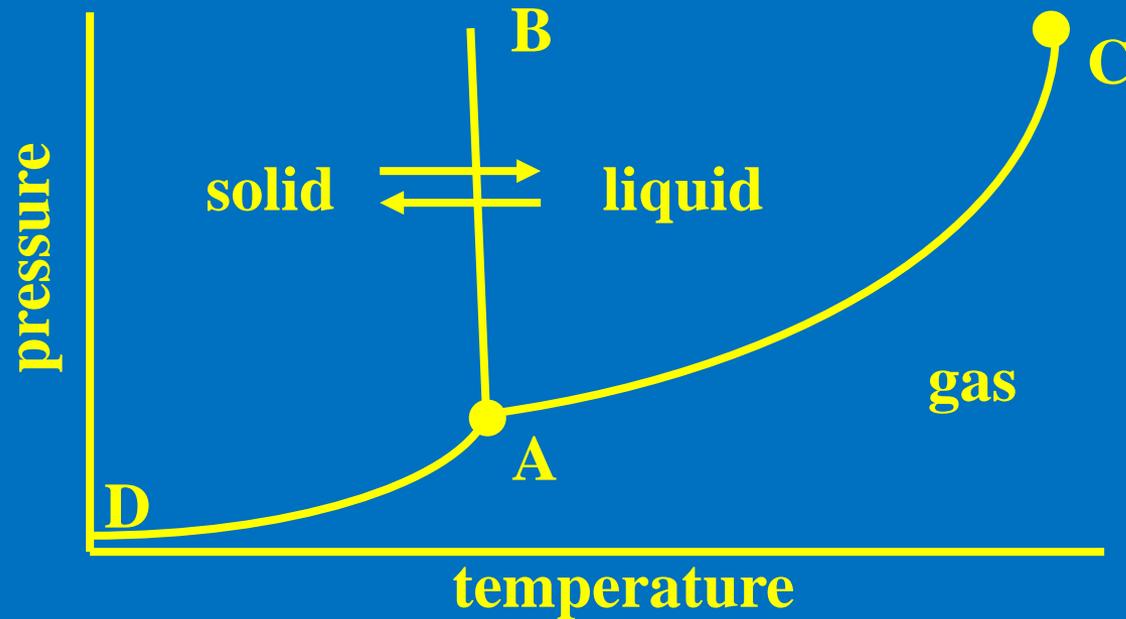
- Typical phase diagram
 - Consists of three curves that divide the diagram into regions labeled:

Solid Liquid Gas



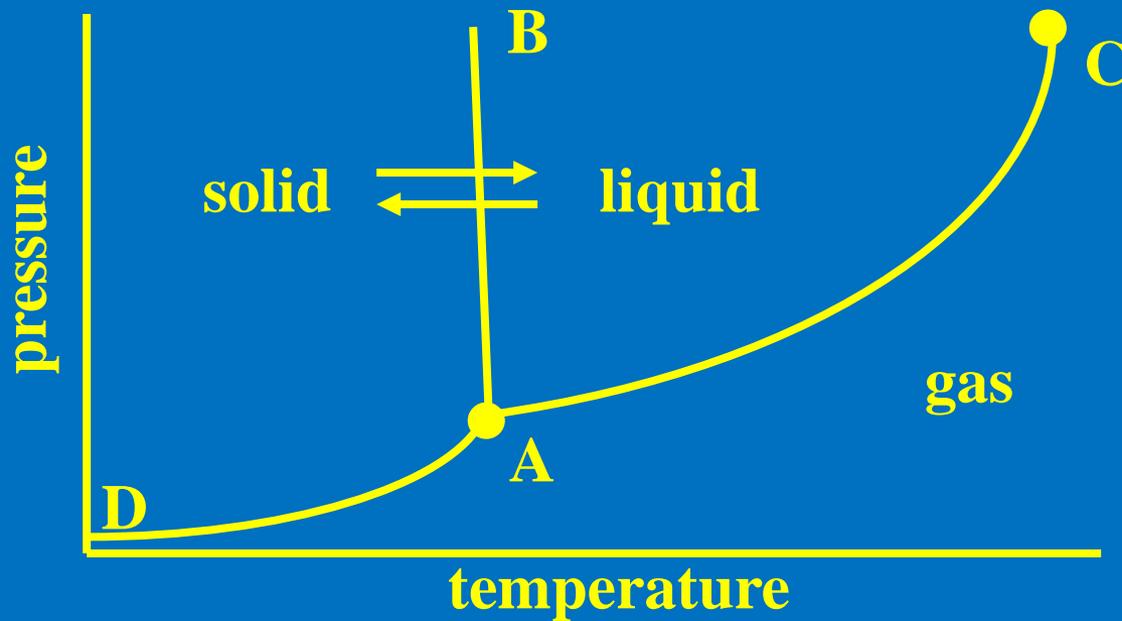
Phase Diagrams

- Curve **AB**, dividing the solid region from the liquid region, represents the conditions under which the solid and liquid are in equilibrium



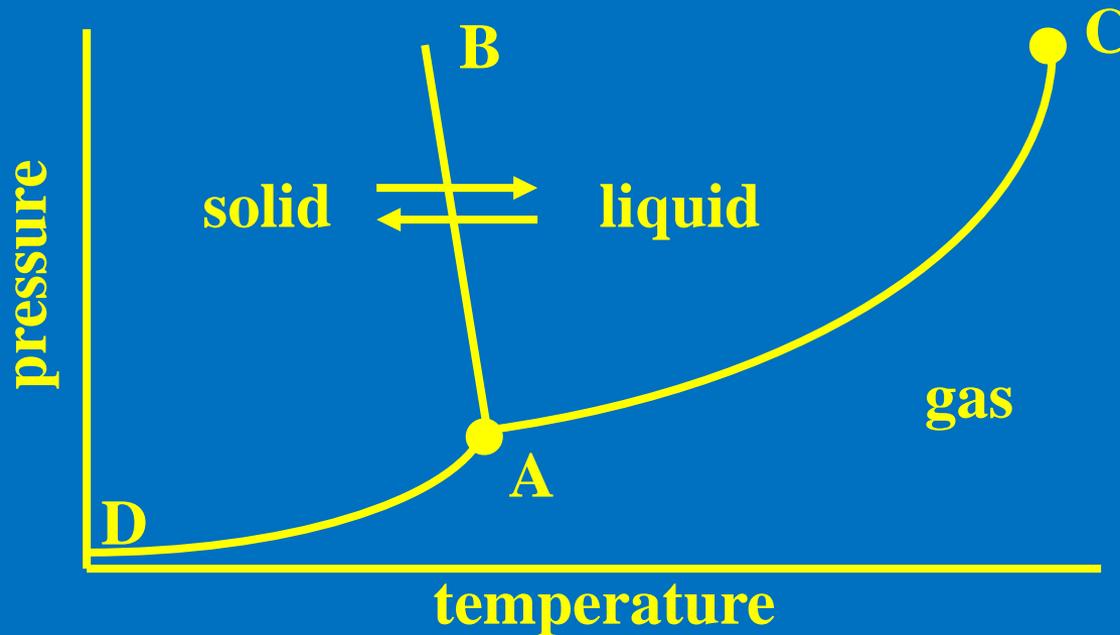
Phase Diagrams

- Usually, the melting point is only slightly affected by pressure. For this reason, the melting point curve, AB, is nearly vertical



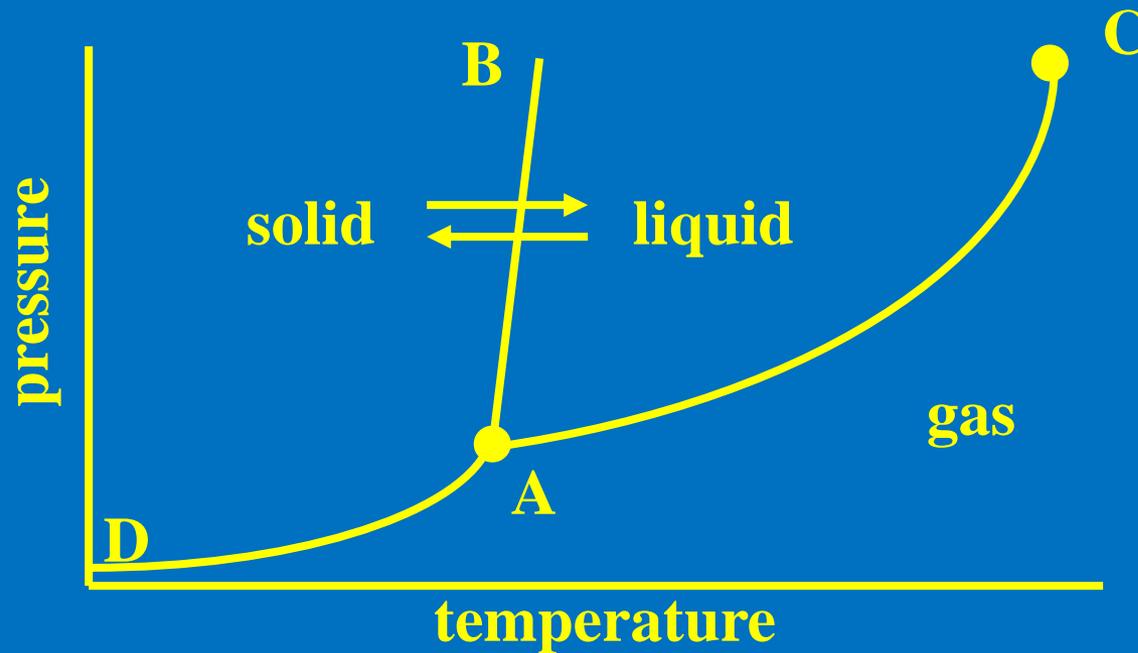
Phase Diagrams

- If a liquid is **more dense** than its solid, the curve leans slightly to the **left**, causing the melting point to **decrease** with pressure



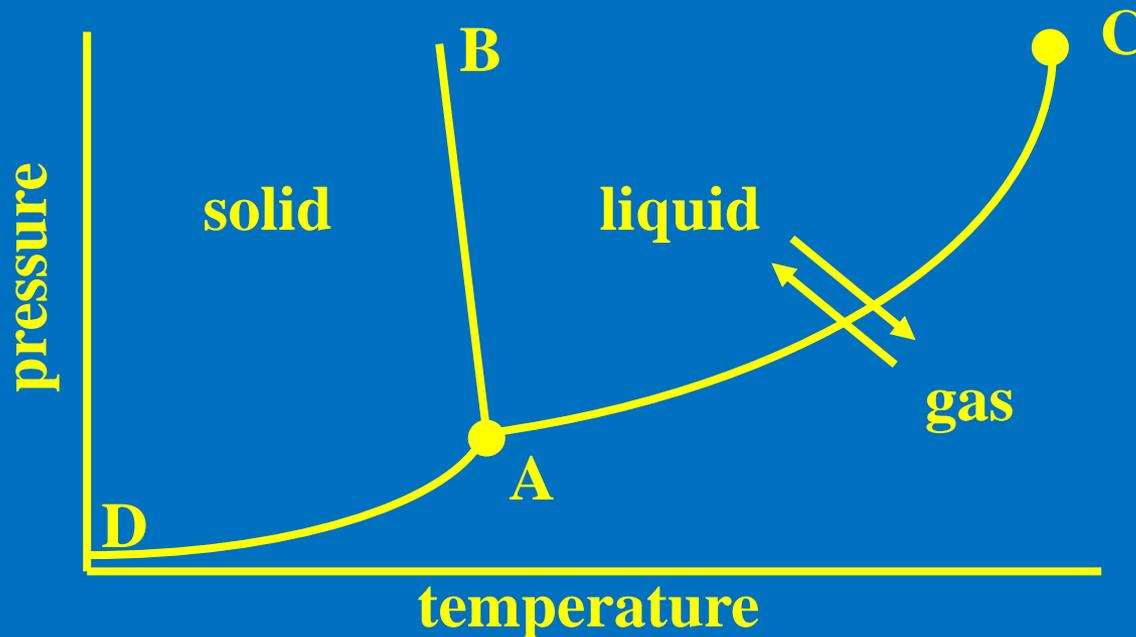
Phase Diagrams

- If a liquid is **less dense** than its solid, the curve leans slightly to the **right**, causing the melting point to **increase** with pressure



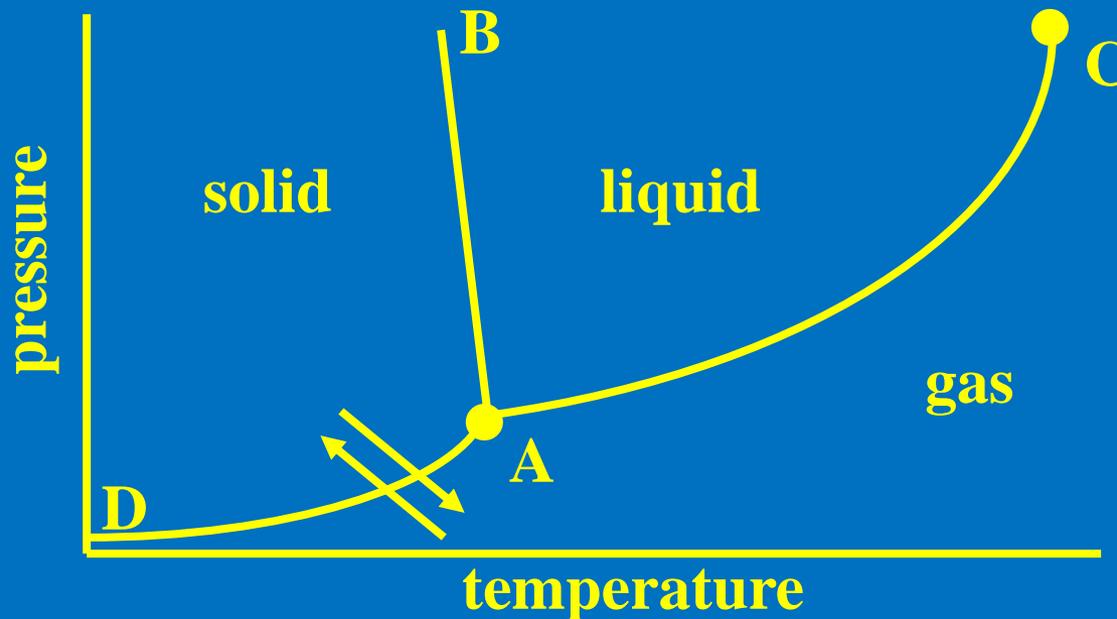
Phase Diagrams

- Curve **AC**, which divides the liquid region from the gaseous region, represents the boiling points of the liquid for various pressures



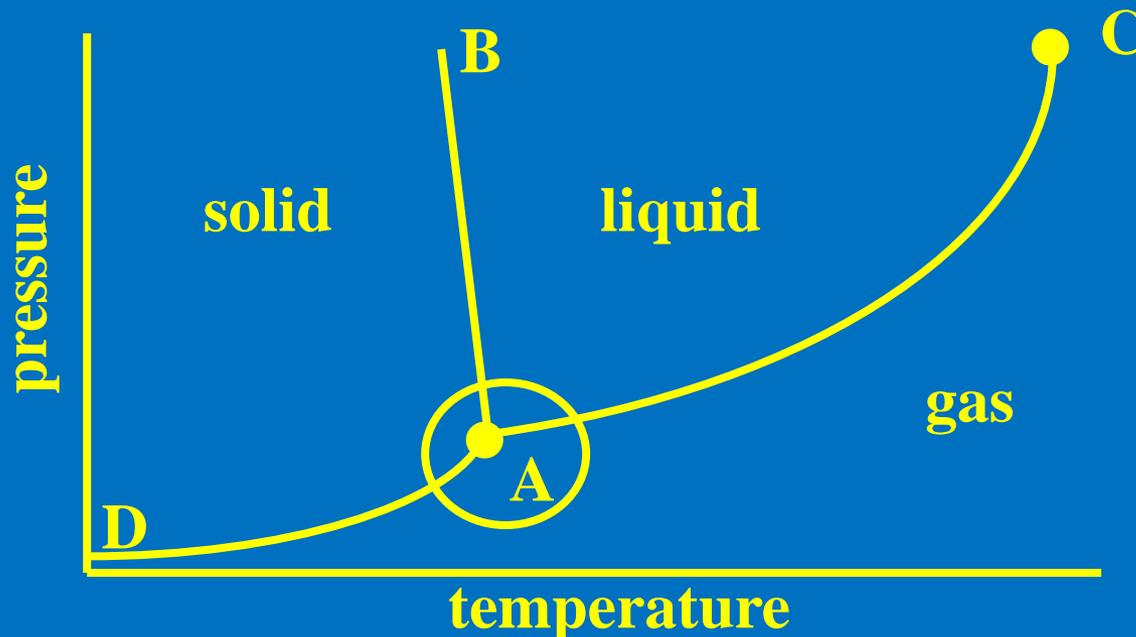
Phase Diagrams

- Curve **AD**, which divides the solid region from the gaseous region, represents the vapor pressures of the solid at various temperatures
- The Phase change of a solid directly to a gas is referred to as "Sublimation"



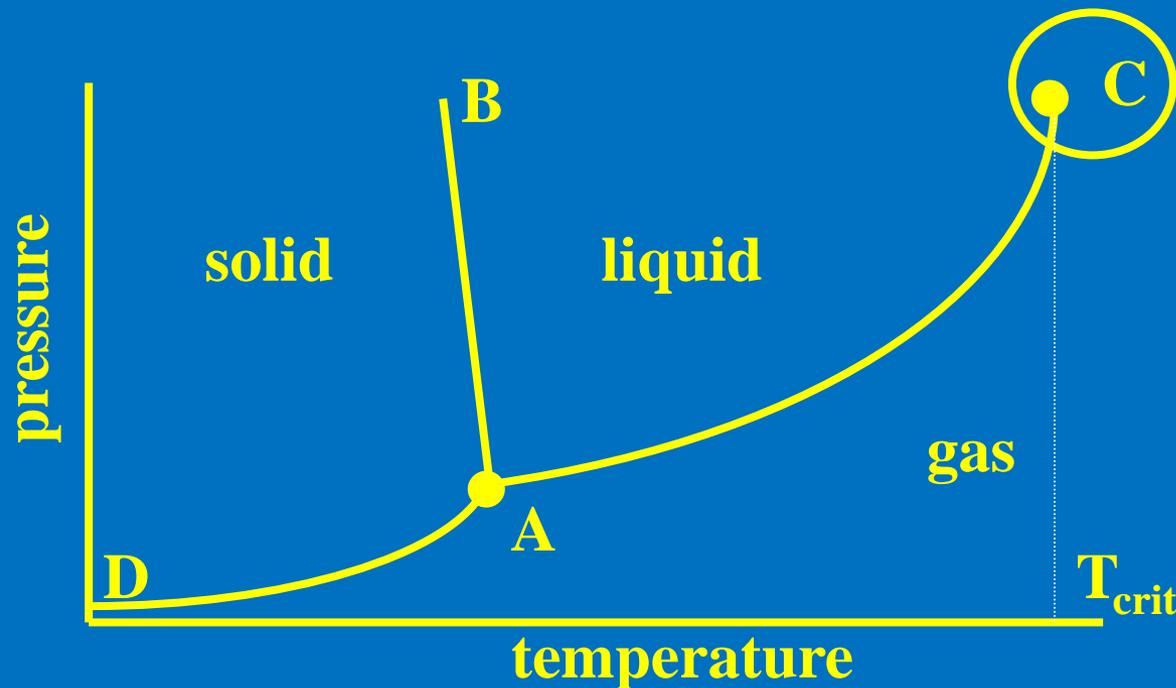
Phase Diagrams

- The curves intersect at **A**, the **triple point**, which is the temperature and pressure where three phases of a substance exist in equilibrium



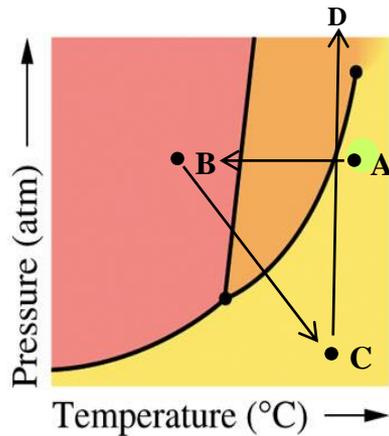
Phase Diagrams

- The temperature above which the liquid state of a substance no longer exists regardless of pressure is called the **critical temperature**



Practice Problem

Shown here is the phase diagram for compound X. The triple point of X is -25.1°C at 0.50 atm and the critical point is 22°C and 21.3 atm .



- What is the state of X at position A?
- If we decrease the temperature from the compound at position A to -28.2°C while holding the pressure constant, what is the state of X?
- If we take the compound starting under the conditions of part b and increase the temperature to 15.3°C and decrease the pressure to 0.002 atm , what is the state of X?
- Would it be possible to make the compound starting under the conditions of part c a solid by increasing just the pressure?

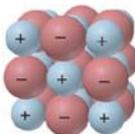
- Position A – Gas
- Position B – Solid
- Position C – Gas
- Position D – Since the solid – liquid line angles to the right, eventually the compound would form a solid (high pressure)

Bonding Forces vs. Non-Bonding Forces

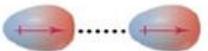
- Intramolecular Forces (Bonding within molecules)
 - Relatively strong because charges are closer together
 - Attractions between:
 - Anions & Cations (Ionic Bonding within molecules)
 - Nuclei & Electron pairs (Covalent Bonding)
 - Metal cations and Delocalized electrons (Metallic)
- Intermolecular Forces (Non-Bonding between molecules)
 - Relatively Weak because charges are smaller and farther apart
 - Phase and phase changes (solid, liquid, gas) are a function of the Intermolecular forces
 - Attractions between:
 - Molecules with partial charges or
 - Molecules and Ions

Bonding Forces vs. Non-Bonding Forces

Intramolecular
(Bonding Forces
within molecules)

Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Bonding				
Ionic		Cation–anion	400–4000	NaCl
Covalent		Nuclei–shared e ⁻ pair	150–1100	H–H
Metallic		Cations–delocalized electrons	75–1000	Fe

Intermolecular
(Non- Bonding Forces
between molecules)

Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Nonbonding (Intermolecular)				
Ion-dipole		Ion charge–dipole charge	40–600	$\text{Na}^+ \cdots \text{O} \begin{matrix} \text{H} \\ \\ \text{H} \end{matrix}$
H bond	$\delta^- \quad \delta^+ \quad \delta^-$ –A–H·····:B–	Polar bond to H–dipole charge (high EN of N, O, F)	10–40	$\begin{matrix} \text{:}\ddot{\text{O}}\text{--H} \\ \\ \text{H} \end{matrix} \cdots \begin{matrix} \text{:}\ddot{\text{O}}\text{--H} \\ \\ \text{H} \end{matrix}$
Dipole-dipole		Dipole charges	5–25	I–Cl·····I–Cl
Ion-induced dipole		Ion charge–polarizable e ⁻ cloud	3–15	$\text{Fe}^{2+} \cdots \text{O}_2$
Dipole-induced dipole		Dipole charge–polarizable e ⁻ cloud	2–10	H–Cl·····Cl–Cl
Dispersion (London)		Polarizable e ⁻ clouds	0.05–40	F–F·····F–F

Intermolecular Forces

- Many of the physical properties of liquids (and certain solids) can be explained in terms of the forces of attraction between molecules, that is, the

Intermolecular Forces

- Three types of forces are known to exist between neutral molecules

- London (or dispersion) forces – weakest
(non-polar molecules)
- Dipole-Dipole forces – stronger
(polar molecules)
- Hydrogen bonding – strongest

(in substances where Hydrogen is directly bonded to either Oxygen, Nitrogen or Fluorine)

London Forces

- **London forces** are the weak attractive forces resulting from **instantaneous dipoles** that occur due to the **distortion of the electron cloud** surrounding a molecule

- London forces:

 - increase with molecular weight**

- The **larger** a molecule, the more easily the electron cloud can be distorted to give an instantaneous dipole

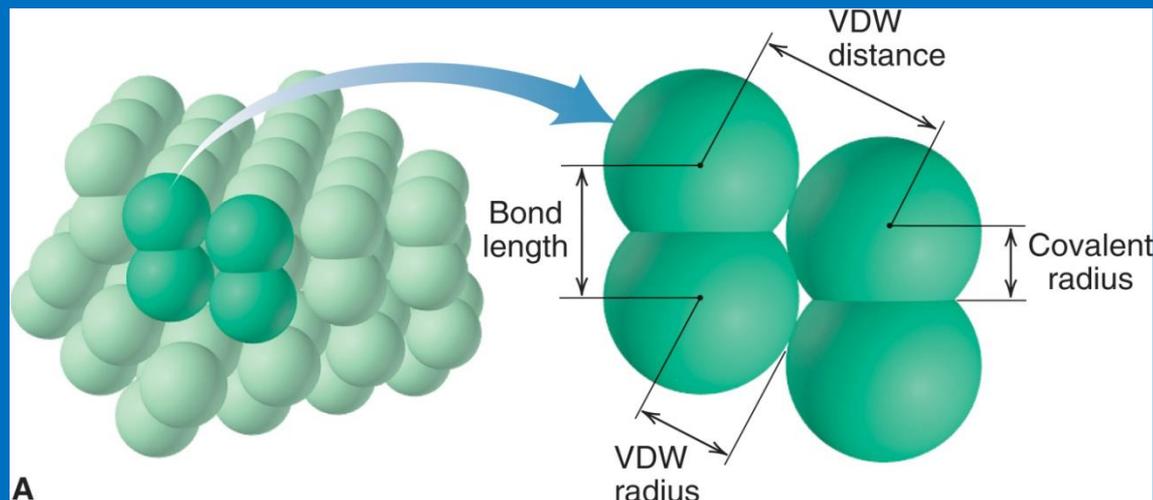
- **All** covalent molecules exhibit some London force

Intermolecular Forces

- **Van der Waals (VDW) forces** – general term that includes

London and Dipole-Dipole forces

- **Van der Waals forces** are the weak attractive forces in a large number of substances
 - The distance between two non-bonded atoms in “**adjacent**” molecules
 - London forces are dispersion forces and are the weakest of the intermolecular forces



Intermolecular Forces

■ Van der Waals (vdw) Forces

➤ Dipole – Dipole

- An external electric field orients gaseous polar molecules
- The polar molecules in liquids and solids lie near each other and their partial charges act as tiny electric fields and give rise to

“Dipole-Dipole” forces

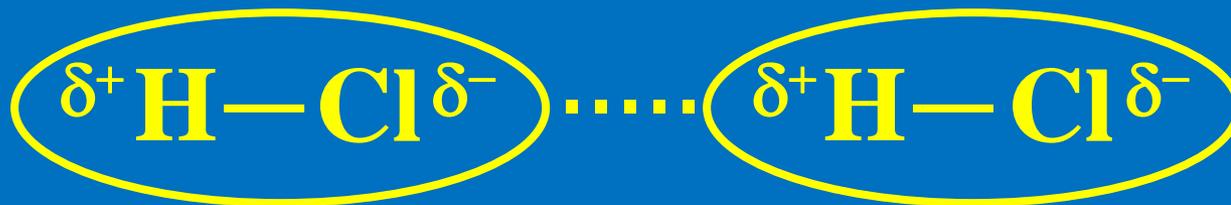
- Dipole-Dipole forces – slightly stronger than the London forces, but less than Hydrogen Bonding
- Magnitude of the Dipole-Dipole force depends on the magnitude of the molecular dipole moment

Dipole-Dipole Forces

- Polar molecules can attract one another through

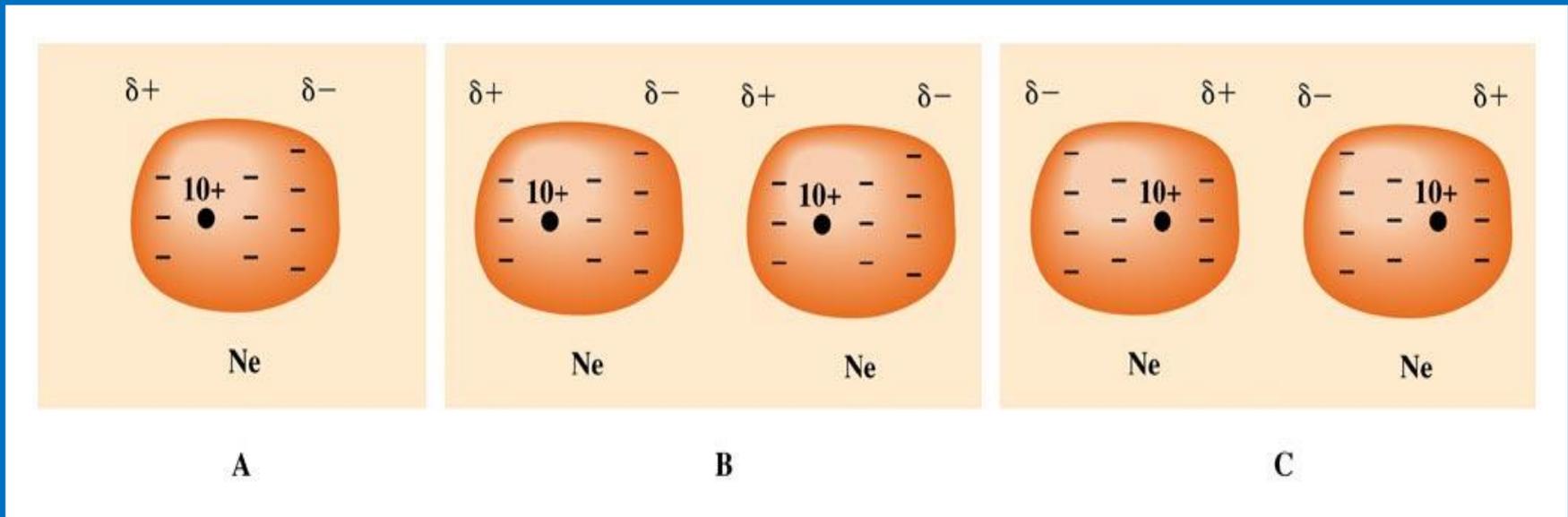
Dipole-Dipole forces

- The **dipole-dipole force** is an attractive intermolecular force resulting from the tendency of **polar molecules** to align themselves positive end to negative end
- Polar molecules arise from the differences in the electronegativity of atoms within a molecule



Origin of the London Force

- At any instant there are more e^- on one side of the Neon (Ne) atom than the other (A)



- Opposite ends of instantaneous dipoles attract (B)
- Electrons on adjacent atoms tend to move together to create new instantaneous dipoles (C)
- **All** molecular compounds have London attractive forces

Intermolecular Forces

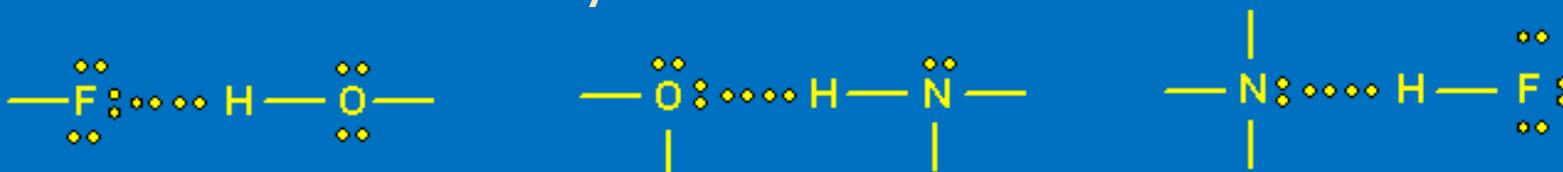
- Polarizability and Induced Dipole Forces
 - A nearby electric field can “induce” a distortion in the electron cloud:
 - Pulling electron density toward a positive pole of the field or
 - Pushing it away from a negative pole
 - For a “nonpolar” molecule the distortion induces a temporary dipole moment
 - For a “polar” molecular, the distortion enhances the dipole moment already present
 - Polarizability increases “Down” a group as size increases (electron cloud is easier to distort)
 - Polarizability decreases across a period because increasing Z_{eff} makes the atoms smaller holding electrons closer

Hydrogen Bonding

- **Hydrogen Bonding** is a special dipole-dipole force that exists between a Hydrogen atom covalently bonded to a very electronegative atom with a lone pair of electrons
 - Only N, O, and F are electronegative enough to leave the Hydrogen nucleus exposed
 - To exhibit Hydrogen Bonding, one of the following three structures must be present



- Hydrogen Bonding between molecules
Note: Hydrogen Bonding between molecules is indicated by dotted line ••••



Sample Problem

- Which of the following substances exhibits Hydrogen Bonding?

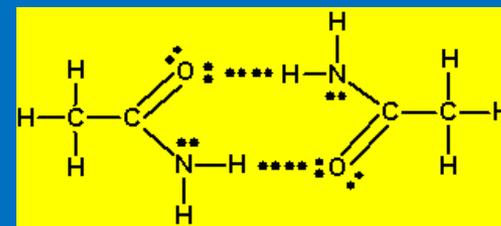
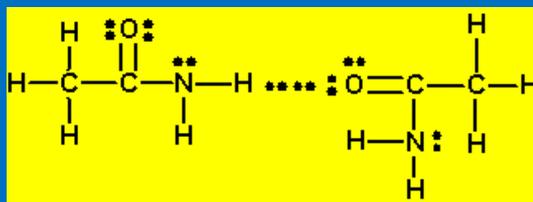
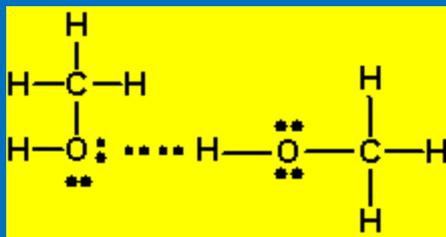
For those that do, draw two molecules of the substance with the H bond(s) between them

- a. C_2H_6 b. CH_3OH c. $CH_3C - NH_2$

Ans: a. C_2H_6 does not form any Hydrogen Bonds

b. For CH_3OH . The **H** covalently bonded to the **O** in one molecule forms an **H-Bond** to the lone pair of electrons on the **O** of an adjacent molecule

c. For $CH_3C - NH_2$. Two of these molecules can form one **H** bond between an **H** bonded to N and the **O**, or they can form two such bonds



Hydrogen Bonding

■ Hydrogen Bonding and Boiling Point

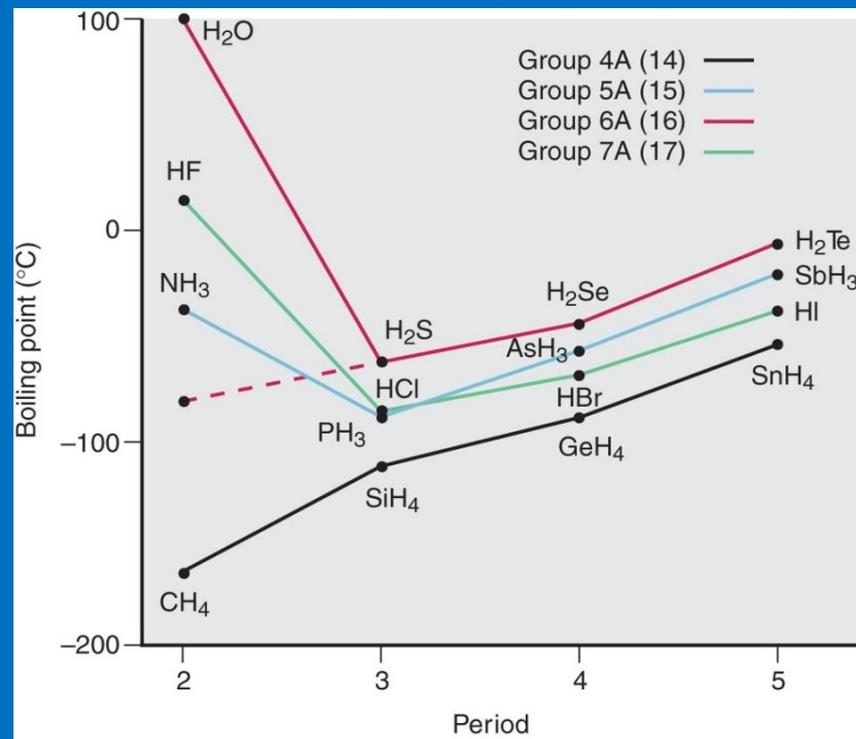
- Boiling points rise as molar mass increases in Group 4A Hydrides



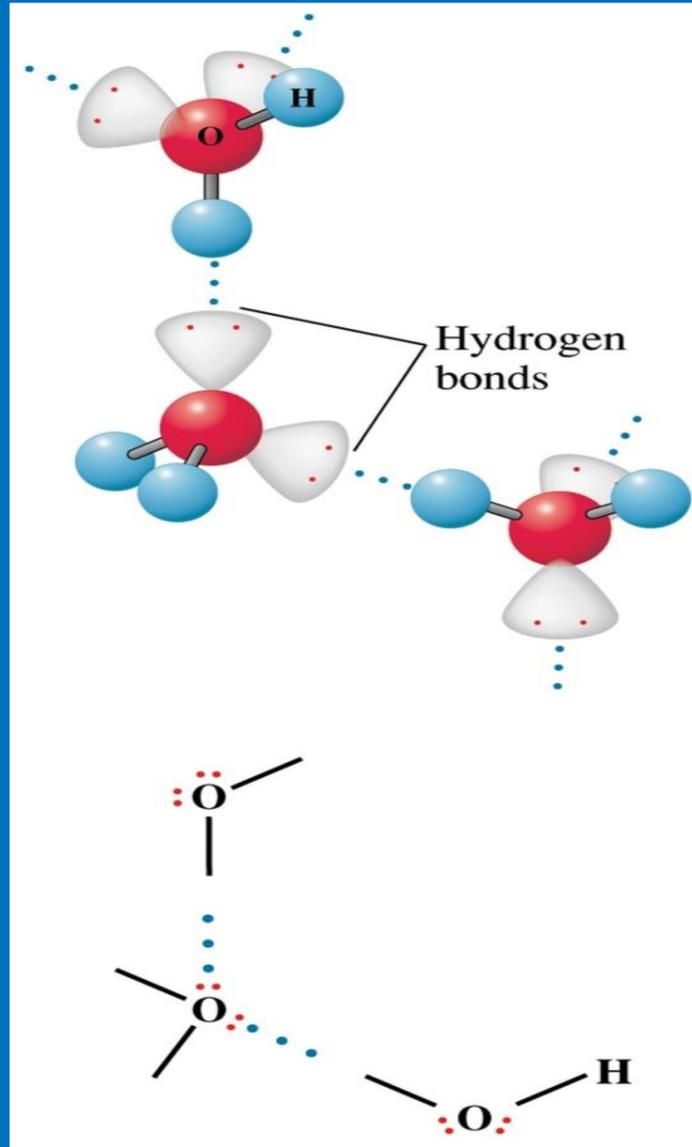
- In the other groups, the first member of each series



The H bonds in these substances require additional energy to break before the molecules can separate and enter the gas phase



Hydrogen Bonding in Water

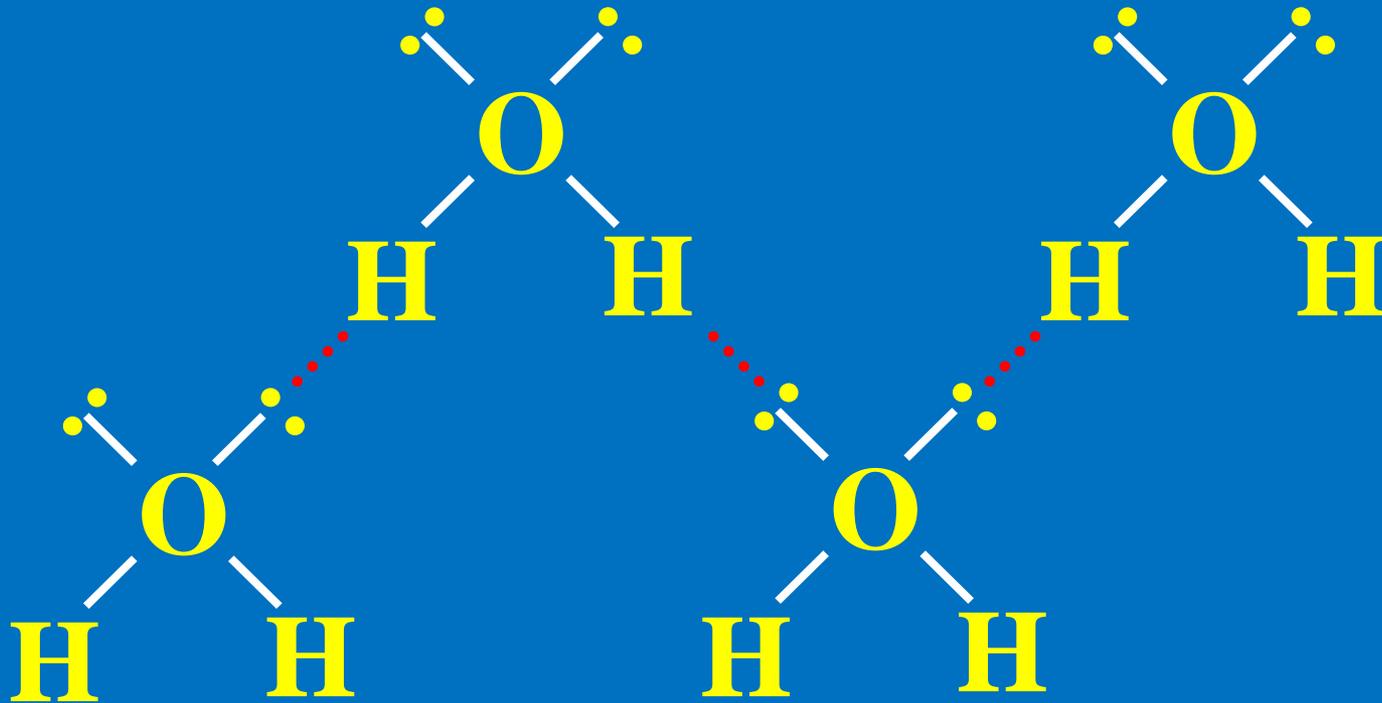


Hydrogen Bonding

- A Hydrogen atom bonded to an electronegative atom appears to be special
 - The electrons in the O-H bond are drawn to the O atom, leaving the dense positive charge of the hydrogen nucleus exposed
 - It's the strong attraction of this exposed nucleus for the lone pair on an adjacent molecule that accounts for the strong attraction
 - A similar mechanism explains the attractions in HF and NH₃

Hydrogen Bonding

Water (H₂O) Molecule

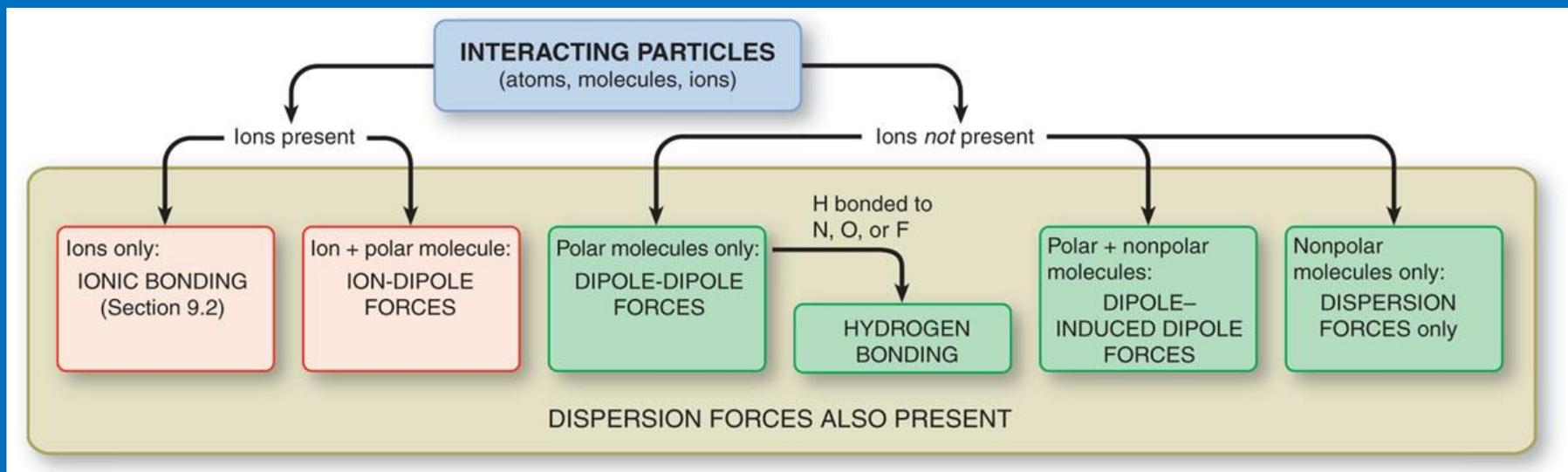


Intermolecular Forces

■ Summary of Intermolecular Forces

- Ions – Ionic Bonding (strongest)
- Polar Molecules – Ion-Dipole; Dipole-Dipole, Hydrogen Bonding
- Polar & Non-Polar – Dipole-Induced Dipole
- Non-Polar – (London) Dispersion (weakest)

■ Note: Dispersion forces are present in ALL molecules



Sample Problem

The forces of attraction between molecules of I_2 are

- a. induced dipole–dipole attractions
- b. dipole–dipole attractions
- c. covalent bonds
- d. London forces
- e. dipole-induced dipole attractions

Ans: d

The I_2 molecule is non-polar

no dipole-dipole no Hydrogen bonding

The forces of attractions between molecules of I_2 are mostly dispersive (London forces)

Sample Problem

How many of the following compounds will exhibit Hydrogen Bonding?



a. 1

b. 2

c. 3

d. 4

e. 5

Ans: 2



H bonded to Nitrogen (N) & Fluorine (F)

Sample Problem

What is the strongest inter-particle force in each substance?



Ans: a. Hydrogen Bonding

b. Non-polar compound – Dispersion forces
(The 4 Cl atoms are positioned symmetrically around the Carbon atom)

c. Non-polar compound – Dispersion forces

Sample Problem

What is the strongest inter-particle force in each substance?

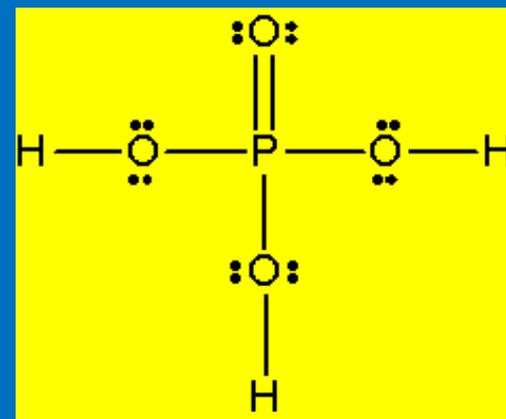


Ans: a. Hydrogen Bonding

b. Dipole-Dipole
(Covalent Molecule)

c. Ionic Bond

(Metal/nonmetal forms ionic compound)



Sample Problem

Which member of each pair of compounds forms intermolecular hydrogen bonds?

a. $\text{CH}_3\text{CHOHCH}_3$ or b. CH_3SCH_3

Ans: a The presence of OH group leads to formation of Hydrogen bonds

a. HF or b. HBr

Ans: a The presence of H attached to highly electronegative F sets up hydrogen bonds

The H in HBr does not form Hydrogen bonds (Br not electronegative enough)

Van der Waals Forces and the Properties of Liquids

- In summary, Van der Waals intermolecular forces:
 - **London (or dispersion) forces (weakest)**
 - **Dipole-dipole forces**

play a large role in many of the physical properties of liquids and gases. These include:

vapor pressure
surface tension

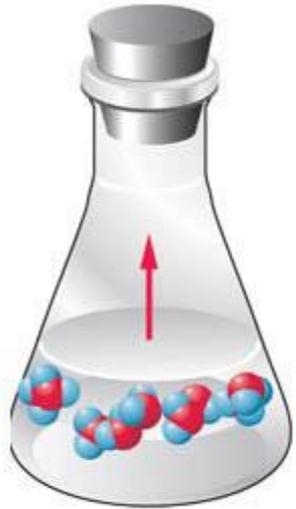
boiling point
viscosity

Vapor Pressure

- Liquids are continuously vaporizing
 - If a liquid is in a closed vessel with space above it, a partial pressure of the vapor state builds up in this space
 - The **vapor pressure** of a liquid is the partial pressure of the vapor over the liquid, measured at equilibrium at a given temperature
 - Raising the temperature of a liquid increases the fraction of molecules moving fast enough to escape the liquid and decreases the fraction moving slowly enough to be recaptured

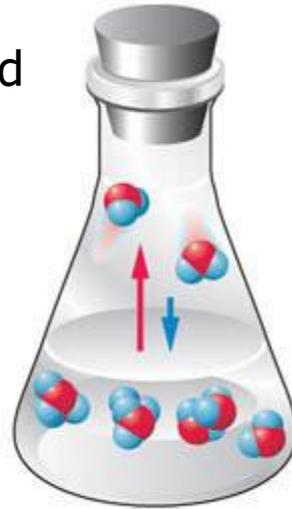
The higher the temperature, the higher the vapor pressure

Rates of Vaporization and Condensation of a Liquid Over Time

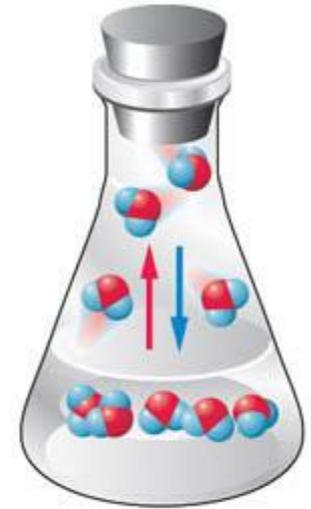


A

Molecules in liquid are in constant motion. Some break free of intermolecular bonds and enter vapor phase



B



C

In C, system comes to equilibrium where **rate of condensation = rate of vaporization**
At this point the vapor pressure is reached

Vapor Pressure

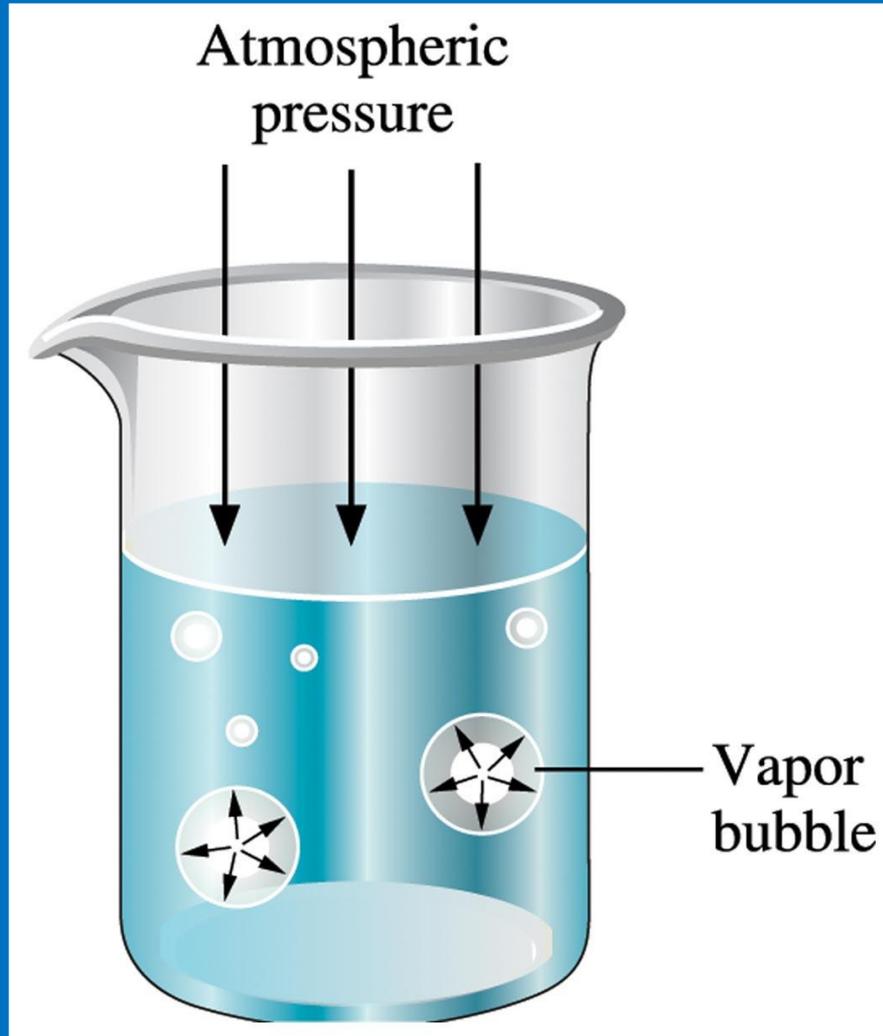
- The vapor pressure of a liquid **depends on its temperature**
 - As the temperature increases, the kinetic energy of the molecular motion becomes greater, and vapor pressure increases
 - Liquids and solids with relatively high vapor pressures at normal temperatures are said to be **volatile**

Boiling Point

- The temperature at which the **vapor pressure** of a liquid equals the pressure exerted on the liquid is called the **boiling point**
 - As the temperature of a liquid increases, the vapor pressure increases until it reaches external pressure (usually atmospheric pressure)
 - At this point, stable bubbles of vapor form within the liquid. This is called **boiling**
 - **Normal Boiling Point** is the boiling point at
1 atmosphere (atm) = 760 torr = 760 mm Hg

Boiling Point

- A liquid boils when the VP inside a bubble equals the external P



Heat of Phase Transition

Heat of Vaporization

- To boil a pure substance from its melting point requires an extra boost of energy to overcome **intermolecular forces**
 - The heat needed to boil 1 mol of a pure liquid substance is called the **Heat of Vaporization** and denoted by ΔH_{vap}
 - For ice at the melting point, the heat of vaporization of water is 40.66 kJ/mol



Endothermic reaction requiring energy input

Heat of Vaporization (vapor pressure & boiling point)

- The heat of vaporization (ΔH_{vap}) is inversely proportional to the vapor pressure

$$\Delta H_{\text{vap}} \propto \frac{1}{\text{Vapor Pressure}}$$

The **higher** the vapor pressure (low boiling point) the **lower** the heat of vaporization

- The heat of vaporization (ΔH_{vap}) is proportional to the boiling point

$$\Delta H_{\text{vap}} \propto \text{Boiling Point} \propto \frac{1}{\text{Vapor Pressure}}$$

Liquids with low vapor pressure require more energy to invoke boiling, i.e., they have higher boiling points

Sample Problem

Which compound should have the lowest heat of vaporization?

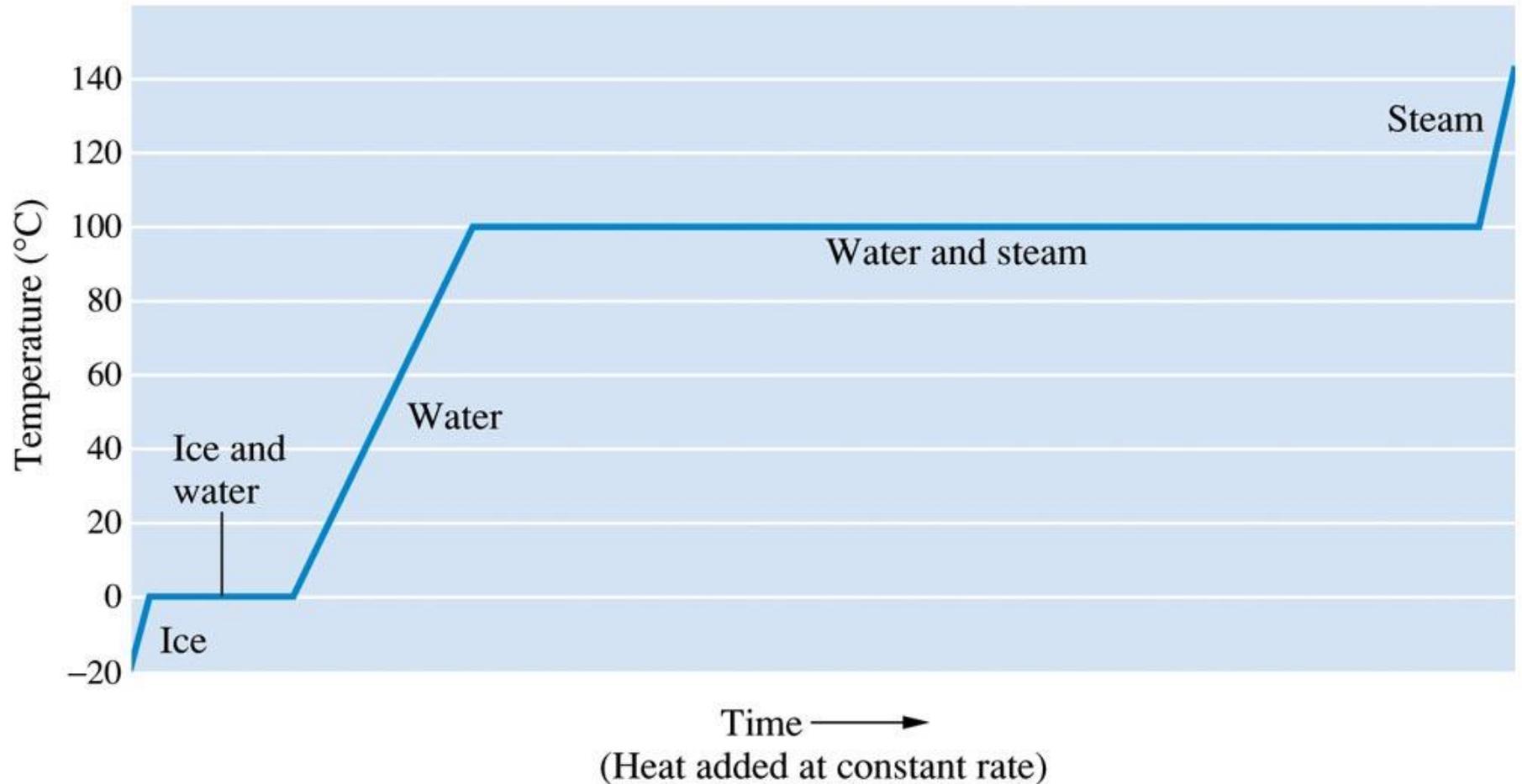


Ans: a

C_5H_{12} has the lowest molecular weight and should have the highest vapor pressure and, thus, the lowest boiling point

Therefore, the Heat of Vaporization (ΔH) should be the lowest

Heating Curve for Water

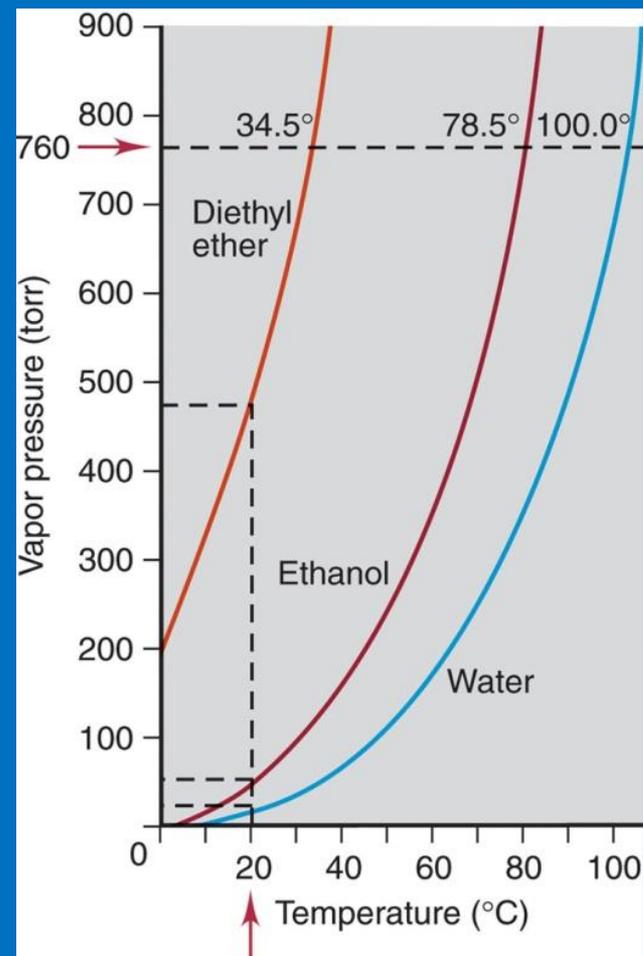
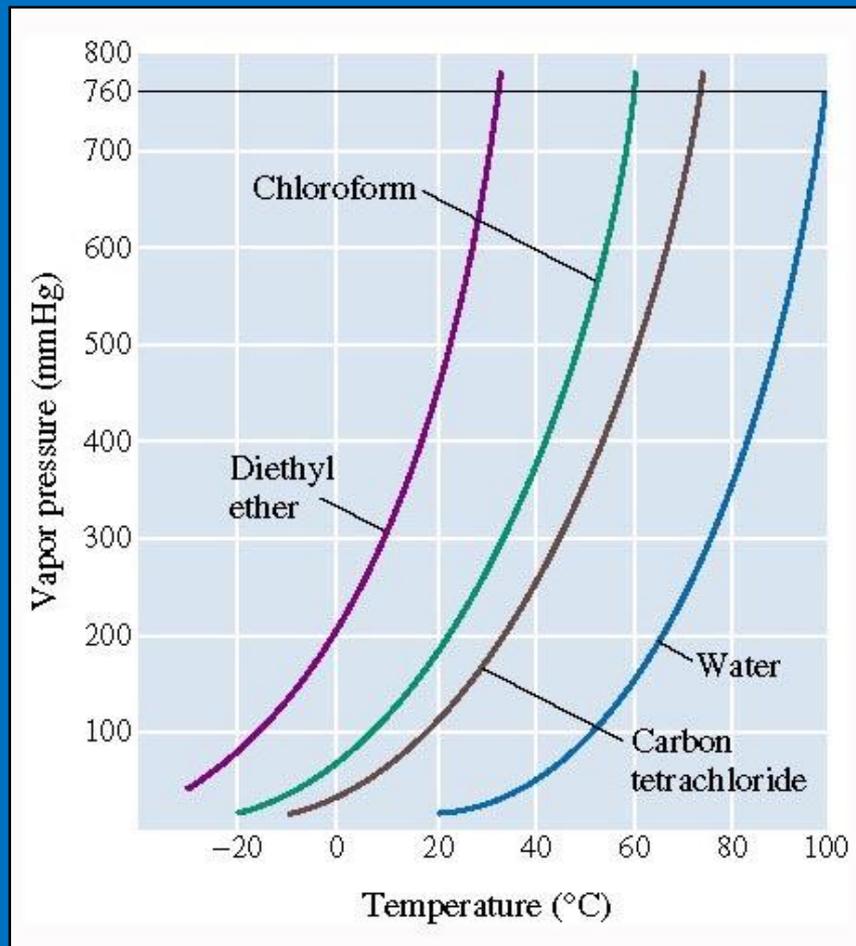


Clausius-Clapeyron Equation

- Recall: Vapor Pressure is a function of Temperature
- Vapor Pressure also depends on the Intermolecular Forces present
- At a given temperature molecules with weaker intermolecular forces have a higher vapor pressure and vaporize more easily
- The weaker the intermolecular forces the higher the vapor pressure

Clausius-Clapeyron Equation

- The impact of the intermolecular forces produces a non-linear relationship between Vapor Pressure and Temperature



Clausius-Clapeyron Equation

- A modification of the Clausius-Clapeyron equation, the two point version, allows for the description of the vapor pressure of a liquid at two different temperatures and pressures

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{T_1 - T_2}{T_2 \times T_1} \right)$$

Note: R can be expressed in 2 different sets of units

$$R = 8.31447 \text{ J/mol} \bullet \text{ K}$$

$$R = 0.0821 \text{ (L} \bullet \text{ atm)/(mol} \bullet \text{ K)}$$

$$\text{Note: } 1 \text{ L} \bullet \text{ atm} = 101.3 \text{ J}$$

Units for Heat of Vaporization - kJ/mol

Sample Problem

The compound Methylchloride, CH_3Cl , has a vapor pressure of 100 mm Hg at 210 K and 10 mm Hg at 181 K.

Calculate the Heat of Vaporization (ΔH_{vap}) of CH_3Cl .

$$P_1 = 100 \text{ mm Hg} \left(\frac{1 \text{ atm}}{760 \text{ mm Hg}} \right) = 0.13158 \text{ atm}$$

$$T_1 = 210^\circ \text{K}$$

$$P_2 = 10 \text{ mm Hg} \left(\frac{1 \text{ atm}}{760 \text{ mm Hg}} \right) = 0.013158 \text{ atm}$$

$$T_2 = 181^\circ \text{K}$$

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{T_1 - T_2}{T_2 \times T_1} \right)$$

$$\Delta H_{\text{vap}} = -\ln \frac{P_2}{P_1} (R) \left(\frac{T_2 \times T_1}{T_1 - T_2} \right)$$

$$\Delta H_{\text{vap}} = -\ln \frac{10}{100} \left(8.31447 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) \left(\frac{181 \times 210}{210 - 181} \text{K} \right)$$

$$\Delta H_{\text{vap}} = (2.30259) \left(8.31447 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (1,311 \text{K})$$

$$\Delta H_{\text{vap}} = 2.51 \times 10^4 \text{ J/mol} = 25.1 \text{ kJ/mol}$$

Sample Problem

Carbon Disulfide, CS_2 , has a heat of vaporization of 29.2 kJ/mol. At 268 K it has a vapor pressure of 100 mmHg. What is its vapor pressure at 301 K?

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{T_1 - T_2}{T_2 \times T_1} \right)$$

$$P_1 = 100 \text{ mm Hg}$$

$$T_1 = 268^\circ \text{ K}$$

$$P_2 = ?$$

$$T_2 = 301^\circ \text{ K}$$

$$\Delta H_{\text{vap}} = 29.2 \text{ kJ/mol} = 29,200 \text{ J/mol}$$

$$R = 8.31447 \text{ J/mol} \cdot \text{K}$$

$$\ln P_2 - \ln P_1 = \left(\frac{-\Delta H_{\text{vap}}}{R} \right) \left(\frac{T_1 - T_2}{T_2 \times T_1} \right)$$

$$\ln P_2 = \left(\frac{-\Delta H_{\text{vap}}}{R} \right) \left(\frac{T_1 - T_2}{T_2 \times T_1} \right) + \ln P_1$$

$$\ln P_2 = \left(\frac{-29,200 \frac{\text{J}}{\text{mol}}}{8.31447 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \right) \left(\frac{268 - 301}{301 \times 268} \text{ K} \right) + \ln(100) \text{ mmHg}$$

$$\ln P_2 = 1.43668 + 4.60517 = 6.04185$$

$$P_2 = 421 \text{ mm Hg}$$

Trouton's Rule

(Heat of Vaporization – Organic liquids)

- The Heat of Vaporization (ΔH_{vap}) of an **organic liquid** can be estimated using Trouton's Rule:

$$\Delta H_{\text{vap}} = 88 \frac{\text{Joules}}{\text{mol} \cdot ^\circ \text{K}} \times T_b (\text{K}) \quad [T_b = \text{Boiling Point}(^\circ \text{K})]$$

Practice Problem

The liquid Tribromomethane, CHBr_3 , has a vapor pressure of 751 Pa at 25 °C and its Boiling Point is 149.5°C.

What is ΔH_{vap} of CHBr_3 ?

- a. 13.2 kJ/mol b. 29.4 kJ/mol c. 37.2 kJ/mol
d. 59.1 kJ/mol e. 70.8 kJ/mol

$$\Delta H_{\text{vap}} = 88 \left(\frac{\text{J}}{\text{mol} \cdot \text{K}} \right) \times T_b = 88 \left(\frac{\text{J}}{\text{mol} \cdot \text{K}} \right) \times ((149.5^\circ \text{C} + 273.15))$$

$$\Delta H_{\text{vap}} = 3.7 \times 10^4 \text{ J / mol} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right)$$

$$\Delta H_{\text{vap}} = 37.2 \text{ kJ / mol}$$

Practice Problem

The liquid Tribromomethane, CHBr_3 , has a vapor pressure of 751 Pa at 25 °C and its Boiling Point is 149.5°C.

What is the vapor pressure in Pa of CHBr_3 at 10 °C?

- a. 2,190 b. 974 c. 698 d. 339 e. 119

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{T_1 - T_2}{T_2 \times T_1} \right) \quad \ln P_2 = \left(\frac{-\Delta H_{\text{vap}}}{R} \right) \left(\frac{T_1 - T_2}{T_2 \times T_1} \right) + \ln P_1$$

$$\ln P_2 = \left(\frac{-37.2 \text{ kJ/mol} \frac{1000 \text{ J}}{1 \text{ kJ}}}{8.31447 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \right) \left(\frac{(25 + 273.15) - (10 + 273.15)}{(10 + 273.15) \times (25 + 273.15) \text{ K}} \right) + \ln(751 \text{ Pa})$$

$$\ln P_2 = \left(\frac{-37,200 \text{ J}}{8.31447 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \right) \left(\frac{15}{84,421 \text{ K}} \right) + 6.6214$$

$$\ln P_2 = -0.79497 + 6.6214 = 5.8264$$

$$P_2 = 339 \text{ Pa}$$

Note: $\Delta H_{\text{vap}} = 37.2 \text{ kJ/mol}$
from previous example

Freezing (Melting) Point

- The temperature at which a pure liquid changes to a crystalline solid, or **freezes**, is called the **freezing point**
 - The **Melting Point** is **identical** to the **Freezing Point** and is defined as the temperature at which a solid becomes a liquid
 - Unlike boiling points, melting points are affected significantly by only large pressure changes

Freezing (Melting) Point Heat of Phase Transition

- To melt a pure substance at its melting point requires an extra boost of energy to overcome **lattice energies**
 - The heat needed to melt 1 mol of a pure substance is called the **heat of fusion, ΔH_{fus}**
 - For ice, the heat of fusion is 6.01 kJ/mol



Sample Problem

How much energy in Joules (J) does it take to melt 1.0 kg of ice at 0°C? ($\Delta H_{\text{fus}}^{\circ} = 6.01 \text{ kJ/mol}$)

$$1 \text{ kg}_{(\text{ice})} \frac{1000 \text{ g}_{(\text{ice})}}{\text{kg}_{(\text{ice})}} \left(\frac{1 \text{ mol}_{(\text{ice})}}{18.016 \text{ g}_{(\text{ice})}} \right) = 55.506 \text{ mol}_{(\text{ice})}$$

$$\Delta H = 55.506 \text{ mol}_{(\text{ice})} \times \left(\frac{6.01 \text{ kJ}}{\text{mol}_{(\text{ice})}} \right) \left(\frac{1000 \text{ J}}{\text{kJ}} \right) = 3.34 \times 10^5 \text{ J}$$

Sample Problem

The heat of combustion of Methane (CH₄) is shown below. How many grams of Methane does it take to melt 1.0 kg of ice?



$$\Delta H(1\text{kg})_{(\text{ice})} = 3.34 \times 10^5 \text{ J / kg} \left(\frac{1\text{kJ}}{1000\text{J}} \right) = 334 \text{ kJ / kg}_{(\text{ice})}$$

$$\left(\frac{-890.0 \text{ kJ}}{\text{mol}_{\text{CH}_4} \times \frac{16.042 \text{ g}}{\text{mol}_{\text{CH}_4}}} \right) = -55.478 \text{ kJ / g}_{\text{CH}_4}$$

Convert ΔH for 1 mol CH₄ to kJ/g of CH₄

$$\frac{\left(\frac{334 \text{ kJ}}{\text{kg}_{(\text{ice})}} \right)}{55.479 \text{ kJ / g}_{(\text{CH}_4)}} = 6.02 \text{ g}_{(\text{CH}_4)} / \text{kg}_{(\text{ice})}$$

Properties of Liquids

Surface Tension

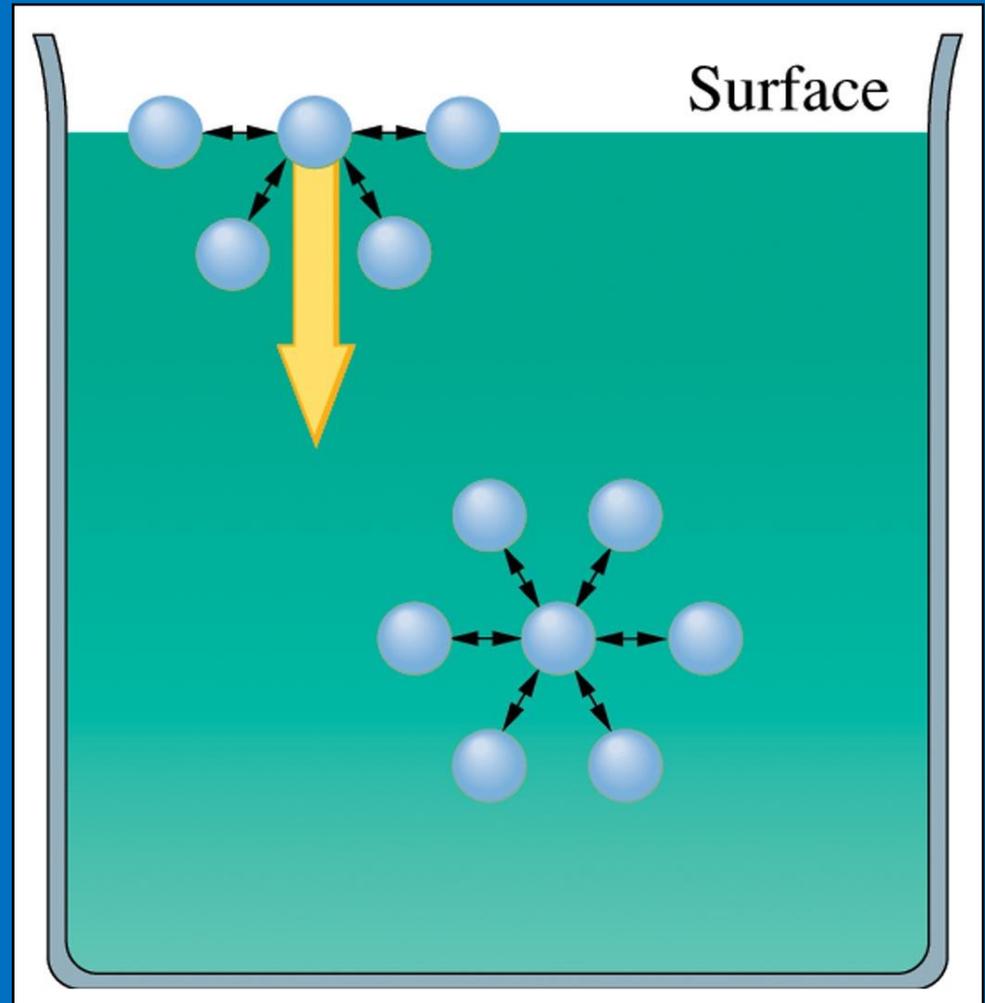
- **Surface tension** is the energy required to increase the surface area of a liquid by a unit amount
 - Surface tension is a contractive tendency of the surface of a liquid that allows it to resist an external force
 - For example, some objects float on the surface of water, even though they are denser than water, e.g. water striders to run on the water surface
 - This property is caused by cohesion of similar molecules, and is responsible for many of the behaviors of liquids

Surface Tension

- A molecule within the bulk of a liquid is pulled equally in all directions by neighboring liquid molecules (cohesion), resulting in a **net force of zero**
- At the surface of the liquid, the molecules are pulled inwards by other molecules deeper inside the liquid but are not attracted as intensely by the molecules in the neighboring medium (be it vacuum, air or another liquid)
- As a result, there is a tendency for the surface area of the liquid to be minimized, i.e. downward pointing meniscus
- This also explains why falling raindrops are nearly spherical, minimizing surface area

Explaining Surface Tension

- Molecule on a liquid surface experiences a net force toward the interior while a molecule in the interior experiences no net force
- The stronger the forces are between the particles in a liquid, the greater the surface tension
- Surfactants (surface-active agents), such as soap and fat emulsifiers, decrease the surface tension of water by congregating at the surface and disrupting the H bonds



Properties of Liquids

Capillarity

- Capillarity – The rising of a liquid through a narrow space against the pull of gravity
- Water
 - In a glass capillary tube (SiO_2), the water molecules form Hydrogen Bonds to the Oxygen atoms of the inner wall of the glass tubes
 - Water will move up the wall of the tube because the adhesive forces from the H-Bonding between the water and the wall are stronger that the cohesive forces (H-bonding) within the water
 - At the same time, the cohesive forces that give rise to surface tension pull the liquid surface taut
 - These adhesive and cohesive forces combine to raise the water level and produces the familiar concave meniscus
 - The liquid rises until gravity pulling down is balanced by the adhesive forces pulling up

Capillarity

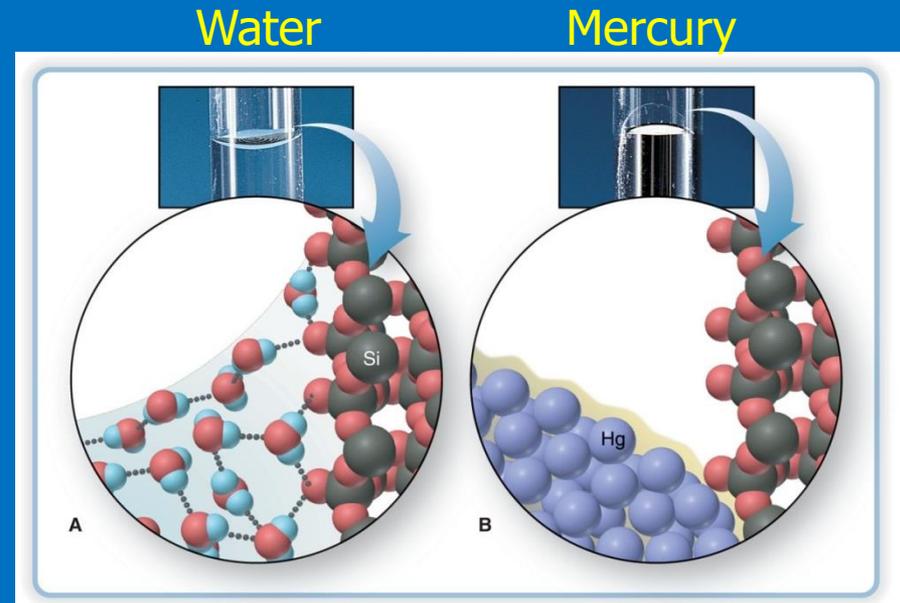
■ Mercury

- The Mercury level in a glass tube immersed in a bowl of Mercury will drop below the level in the bowl
- Mercury has a **higher surface tension** than water because of the **stronger cohesive forces** from metallic bonding
- The cohesive forces among the Mercury atoms are much stronger than the adhesive forces (mostly dispersion) between Mercury and glass, thus the Mercury atoms tend to pull away from the walls of the glass tube
- At the same time, the surface atoms near the glass wall are being pulled toward the interior of the Mercury by its high surface tension forcing the overall level to drop forming a convex (upward pointing) meniscus

Properties of Liquids

Capillarity

- Water displays a concave meniscus in a glass tube
 - Adhesive (H-Bond) forces between H_2O and O-Si-O groups of the glass are stronger than the cohesive forces (H-Bond) within the water
- Mercury displays a convex meniscus in a glass tube
 - The cohesive (metallic bond) forces within mercury are stronger than the adhesive (dispersion) forces between the mercury and the glass



In a glass tube, water molecules are pulled toward the glass surface creeping up the tube, but the surface molecules are being pulled down; combined forces form a downward concave meniscus. Mercury molecules are pulled away from the glass forming an upward (convex) surface curve. The surface atoms are pulled downward dropping the level below the dish level.

Properties of Liquids

Viscosity

- **Viscosity** is a measure of the resistance to flow (units of $\text{kg/m}\cdot\text{s}$) exhibited by all liquids and gases being deformed by either **shear** stress or **extensional** stress
 - Viscosity describes a fluid's internal resistance to flow and may be thought of as a measure of **fluid friction**
 - Viscosity decreases with increase temperature
 - Viscosity can be illustrated by measuring the time required for a steel ball to fall through a column of the liquid
 - Even without such measurements, you know that syrup has a greater viscosity than water

Van der Waals Forces and the Properties of Liquids

- The molecular structure of a substance defines the **intermolecular forces** holding it together
 - Many physical properties of substances are attributed to their intermolecular forces
 - These properties include:

Vapor Pressure

Boiling Point

Surface Tension

Viscosity

Van der Waals Forces and the Properties of Liquids

- **Vapor pressure** (also known as *equilibrium vapor pressure*), is the **pressure** of a **vapor** in **equilibrium** with its **non-vapor phases**
- The atmospheric pressure boiling point of a liquid (also known as the **normal boiling point**) is the temperature where the vapor pressure equals the ambient atmospheric pressure
- The **vapor pressure** of a liquid depends on intermolecular forces
- When the intermolecular forces in a liquid are **strong**, you expect the vapor pressure to be **low**
- The **normal boiling point** is related to vapor pressure and is **lowest** for liquids with the **weakest** intermolecular forces

Van der Waals Forces and the Properties of Liquids

Vapor Pressure Summary:

- The vapor pressure of any substance increases non-linearly with temperature according to the **Clausius-Clapeyron relation**
- The higher the vapor pressure of a liquid at a given temperature, the lower the normal boiling point (i.e., the boiling point at atmospheric pressure) of the liquid

High VP – Weak IM Forces
Low Boiling Point

Low VP – Strong IM Forces
High Boiling Point

Van der Waals Forces and the Properties of Liquids

■ Surface tension

- Intermolecular forces produce attractions between the molecules of a liquid
- Intermolecular forces exert different effects on a molecule at the surface of the liquid than on a molecule in the interior (bulk) of the liquid
- There is a net attraction downward and move toward the interior to increase attractions and become more stable
- As intermolecular forces between molecules increase, the apparent surface tension also increases
- The liquid surface tends to have the smallest possible area (sphere-like)

Van der Waals Forces and the Properties of Liquids

■ Viscosity

- A liquid's viscosity (resistance to flow) results from intermolecular attractions that impede the ability of the molecules to slide around each other
- Both gases and liquids flow, but liquid viscosities are much higher because intermolecular forces operate over much shorter distances
- When molecules move faster at higher temperatures, they can overcome (reduce the effect of) intermolecular forces more easily, thus reducing the resistance to flow
- As intermolecular forces increase, the resistance to flow (viscosity) usually increases

Intermolecular vs Chemical Bonding

Types of Intermolecular and Chemical Bonding Interactions

Type of Interaction	Approximate Energy (kJ/mol)
Intermolecular	
Van der Waals (dipole–dipole, London)	0.1 to 10
Hydrogen bonding	10 to 40
Chemical bonding	
Ionic	100 to 1000
Covalent	100 to 1000

Properties of Liquids

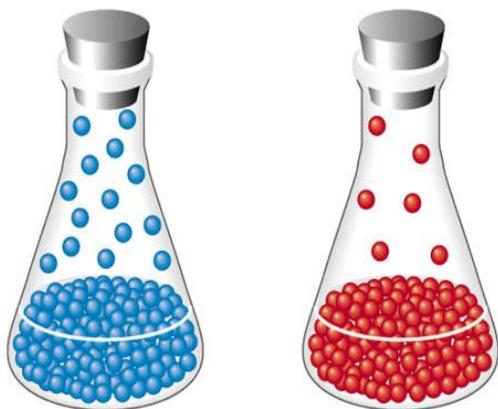
Properties of Some Liquids at 20°C

Substance	Molecular Weight (amu)	Vapor Pressure (mmHg)	Surface Tension (J/m ²)	Viscosity (kg/m · s)
Water, H ₂ O	18	1.8×10^1	7.3×10^{-2}	1.0×10^{-3}
Carbon dioxide, CO ₂	44	4.3×10^4	1.2×10^{-3}	7.1×10^{-5}
Pentane, C ₅ H ₁₂	72	4.4×10^2	1.6×10^{-2}	2.4×10^{-4}
Glycerol, C ₃ H ₈ O ₃	92	1.6×10^{-4}	6.3×10^{-2}	1.5×10^0
Chloroform, CHCl ₃	119	1.7×10^2	2.7×10^{-2}	5.8×10^{-4}
Carbon tetrachloride, CCl ₄	154	8.7×10^1	2.7×10^{-2}	9.7×10^{-4}
Bromoform, CHBr ₃	253	3.9×10^0	4.2×10^{-2}	2.0×10^{-3}

Practice Problem

Conceptual Problem

Consider two flasks that each contains different pure liquids at 20°C.



The liquid in one flask, substance A, has a molar mass of 100 g/mol and has hydrogen bonding. The liquid in the other flask, substance B, has a molar mass of 105 g/mol and has dipole-dipole interactions.

- If the molecular structures of the compounds are very similar, which flask likely contains substance A?
- If you were to increase the temperature of each of the flasks by 15°C, how would the pictures change (assume that you stay below the boiling points of the liquids)?

- The flask with the red molecules has a lower Vapor Pressure; thus a higher boiling point due to "strong" intermolecular forces", probably Hydrogen Bonding.

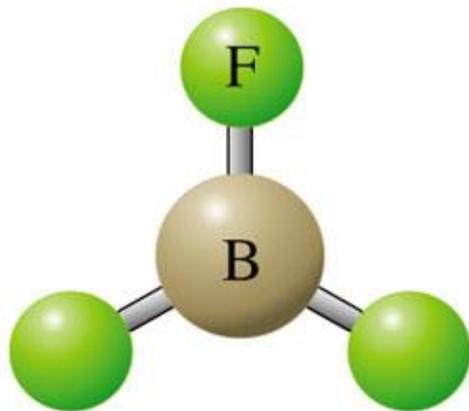
The flask with the blue molecules has a higher vapor pressure and lower boiling point due to weak intermolecular forces – dispersion & dipole-dipole

- Increasing the temperature will increase the vapor pressure of both liquids, but would not change the relative vapor pressures

Practice Problem

For each of the following substances,
List the kinds of intermolecular forces expected

a.

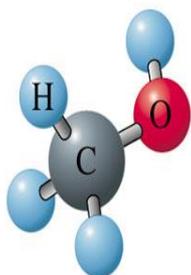


- b. isopropyl alcohol, $\text{CH}_3\text{CHOHCH}_3$
- c. hydrogen iodide, HI
- d. krypton, Kr

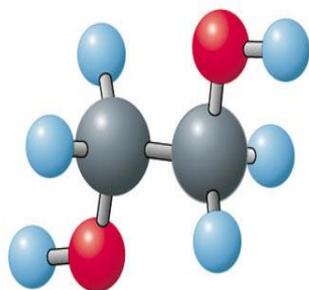
- a. Molecule is non-polar with mostly dispersion forces**
- b. Isopropyl Alcohol is Polar with some dispersion, but mostly Hydrogen Bonding**
- c. Iodine is not electronegative enough to support Hydrogen bonding \therefore dipole-dipole**
- d. Krypton is a monatomic element \therefore dispersive forces**

Practice Problem

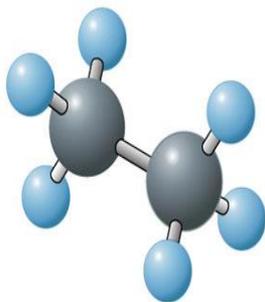
List the following substances in order of increasing boiling point



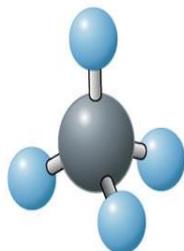
(a)



(b)



(c)



(d)

Increasing Boiling Pt

D methane	–	-161 °C
C ethane	–	- 88.6°C
A methanol	–	64.7°C
B ethylene glycol	–	197.2°C

Methane & Ethane are non-polar with dispersion forces. Increasing Mol Wgt increases Boiling Point.

Methanol and Ethylene Glycol both have hydrogen bonding

Ethylene glycol has more Hydrogen bonding than Methanol

Sample Problem

Which of the following compounds is expected to have the **HIGHEST boiling point?**

- a. CH_3OCH_3 b. $\text{CH}_3\text{CH}_2\text{OH}$ c. $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$
d. $\text{CH}_3\text{CH}_2\text{CH}_3$ e. CH_3Cl

Ans: b (Ethyl Alcohol ($\text{CH}_3\text{CH}_2\text{OH}$) – 78.29 °C)

Ethanol has Hydrogen Bonding

Sample Problem

Which of the following compounds is expected to have the **HIGHEST vapor pressure? Lowest VP?**



Highest Vapor Pressure (lowest boiling point)
(intermolecular dipole-dipole)

Ans: b. Dimethyl Ether (CH_3OCH_3 bp -24.8°C)

Lowest Vapor Pressure (highest boiling point)
(Hydrogen Bonding)

Ans: c. Ethyl Alcohol ($\text{CH}_3\text{CH}_2\text{OH}$) bp 78.29°C

Solid State

- A **solid** is a nearly incompressible state of matter with a well-defined shape
- The units making up the solid are in **close contact** and in **fixed positions**
- Solids are characterized by the **type of force** holding the structural units together
- In some cases, these forces are **intermolecular**, but in others they are **chemical bonds** (metallic, ionic, or covalent)

Solid State

- From this point of view, there are four types of solids

Relative Bond Energy

- Molecular (Van der Waals forces) 0.01 – 660 kJ/mol
- Metallic (Metallic bond) 75 – 1000 kJ/mol
- Covalent (Covalent bond) 150 – 1100 kJ/mol
- Ionic (Ionic bond) 400 – 4000 kJ/mol

Types of Solids

- A **molecular solid** is a solid that consists of atoms or molecules held together by intermolecular forces.
- A **metallic solid** is a solid that consists of positive cores of atoms held together by a surrounding “sea” of electrons (metallic bonding)
- An **ionic solid** is a solid that consists of cations and anions held together by electrical attraction of opposite charges (ionic bond)
- A **covalent network solid** is a solid that consists of atoms held together in large networks or chains by covalent bonds

Physical Properties

- Many physical properties of a solid can be attributed to its structure

Melting Point and Structure

- For a solid to melt, the forces holding the structural units together must be overcome
- For a **molecular solid**, these are weak intermolecular attractions
- Thus, **molecular solids tend to have low melting points** (below 300 °C)

Physical Properties

- Many physical properties of a solid can be attributed to its structure (con't)

Melting Point and Structure

- For **ionic solids** and **covalent network solids** to melt, chemical bonds must be broken
- For that reason, **their melting points are relatively high**
 - Note that for **ionic solids**, melting points increase with the strength of the ionic bond
 - Ionic bonds are stronger when:
 - ◆ The magnitude of charge is high
 - ◆ The ions are small (higher charge density)

Physical Properties

■ Melting Point and Structure

- Metals often have high melting points, but there is considerable variability
- Melting points are low for Groups IA and IIA but increase as you move into the transition metals
- The elements in the middle of the transition metals have the highest melting points

Sample Problem

- Which of the following compounds would be expected to have the highest melting point?



Ans: b

CO_2 , NO_2 , N_2O_5 are covalent gases

P_4O_{10} (mp 560°C) is a crystalline covalent solid

SiO_2 (mp 1650°C) molecules have stronger dipole-dipole intermolecular than P_4O_{10}

Physical Properties

■ **Hardness and Structure**

- **Hardness** depends on how easily structural units can be moved relative to one another
- **Molecular solids** with weak intermolecular attractions are **rather soft** compared with ionic compounds, where forces are much stronger
- **Covalent network solids** are quite hard because of the rigidity of the covalent network structure
- **Molecular and ionic crystals** are generally brittle because they fracture easily along crystal plane
- **Metallic solids**, by contrast, are malleable

Physical Properties

■ **Electrical Conductivity and Structure**

- Molecular and ionic solids are generally considered nonconductors
- Ionic compounds conduct in their molten state, as ions are then free to move
- Metals are all considered conductors
- Of the covalent network solids, only graphite conducts electricity
 - This is due to the delocalization of the resonant p electrons in graphite's sp^2 hybridization

Crystalline Solids

Crystal Lattices and Unit Cells

- Solids can be crystalline or amorphous
 - A **crystalline solid** is composed of one or more crystals; each crystal has a well-defined, ordered structure in three dimensions
 - Examples include sodium chloride and sucrose
 - An **amorphous solid** has a disordered structure. It lacks the well-defined arrangement of basic units found in a crystal
 - Glass is an amorphous solid

Crystal Lattices

- A **crystal lattice** is the geometric arrangement of lattice points in a crystal
 - A **unit cell** is the smallest boxlike unit from which you can construct a crystal by stacking the units in three dimensions
 - There are **seven (7)** basic shapes possible for unit cells, which give rise to **seven crystal systems** used to classify crystals

Crystal Lattices

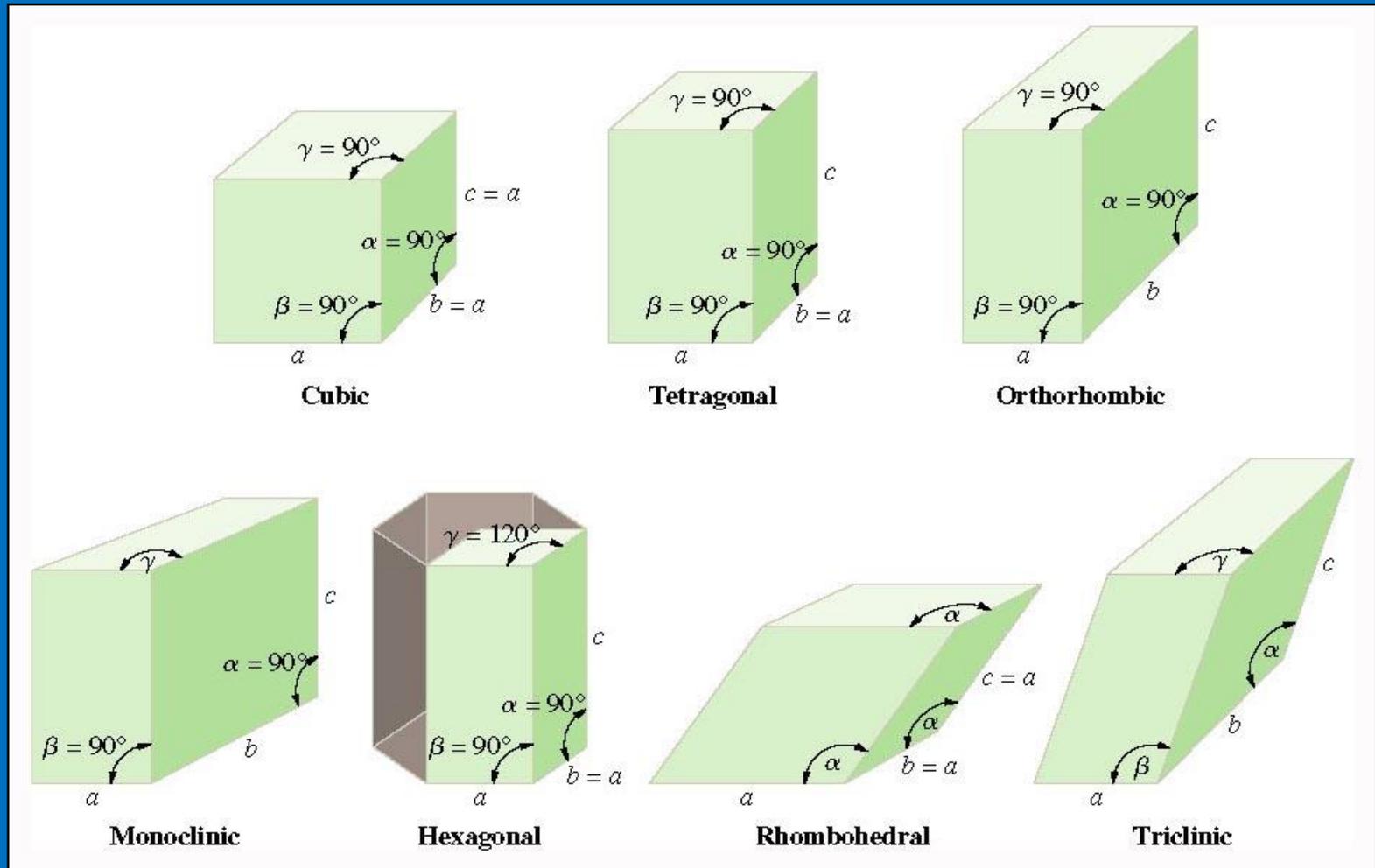
■ Crystal Lattice (Con't)

- These seven systems can have more than one possible crystal lattice
- A “primitive” lattice has lattice points only at the corners of each cell
- Other lattices in the same crystal may have lattice points on the “faces” of the unit cell

Crystal Lattices

■ Crystal Lattice

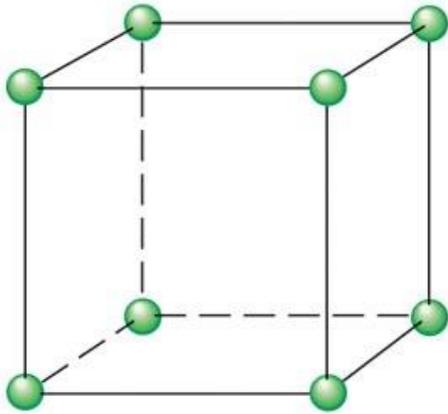
➤ Unit-Cell Shapes of the 7 Different Crystal Systems



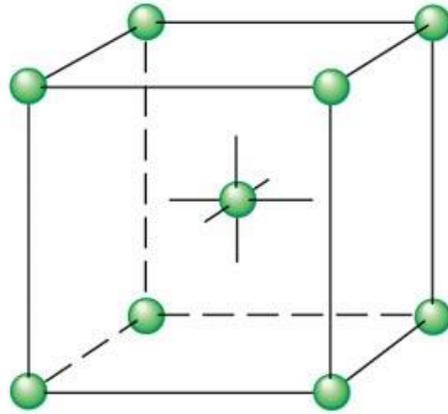
Cubic Unit Cells

- A **simple cubic unit cell** is a cubic cell in which the **lattice** points are situated only at the corners
- A **body-centered cubic unit cell** is one in which there is a **lattice** point in the center of the cell as well as at the corners
- A **face-centered cubic unit cell** is one in which there are **lattice** points at the center of each face of the cell as well as at the corners

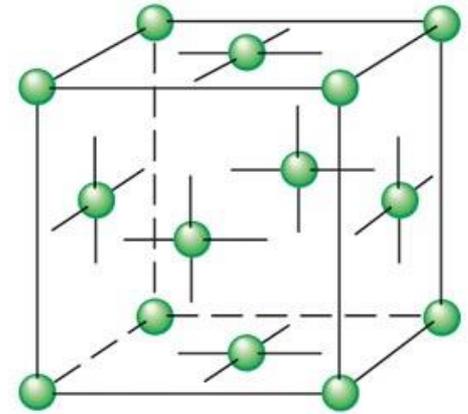
Cubic Unit Cells



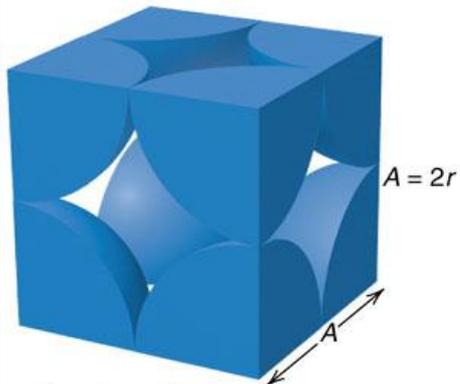
Simple cubic



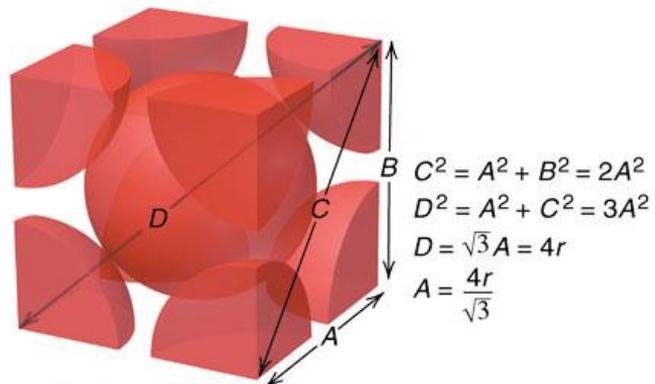
Body-centered cubic



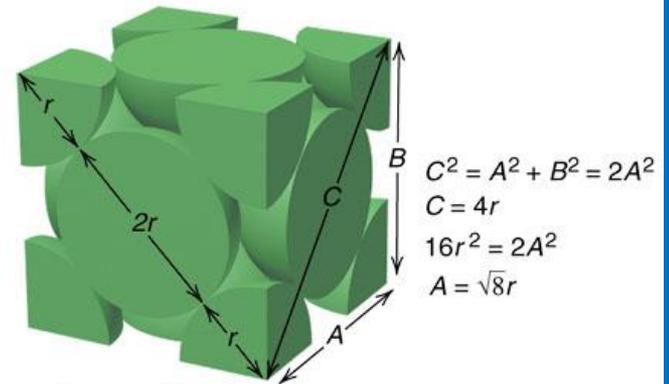
Face-centered cubic



Simple cubic



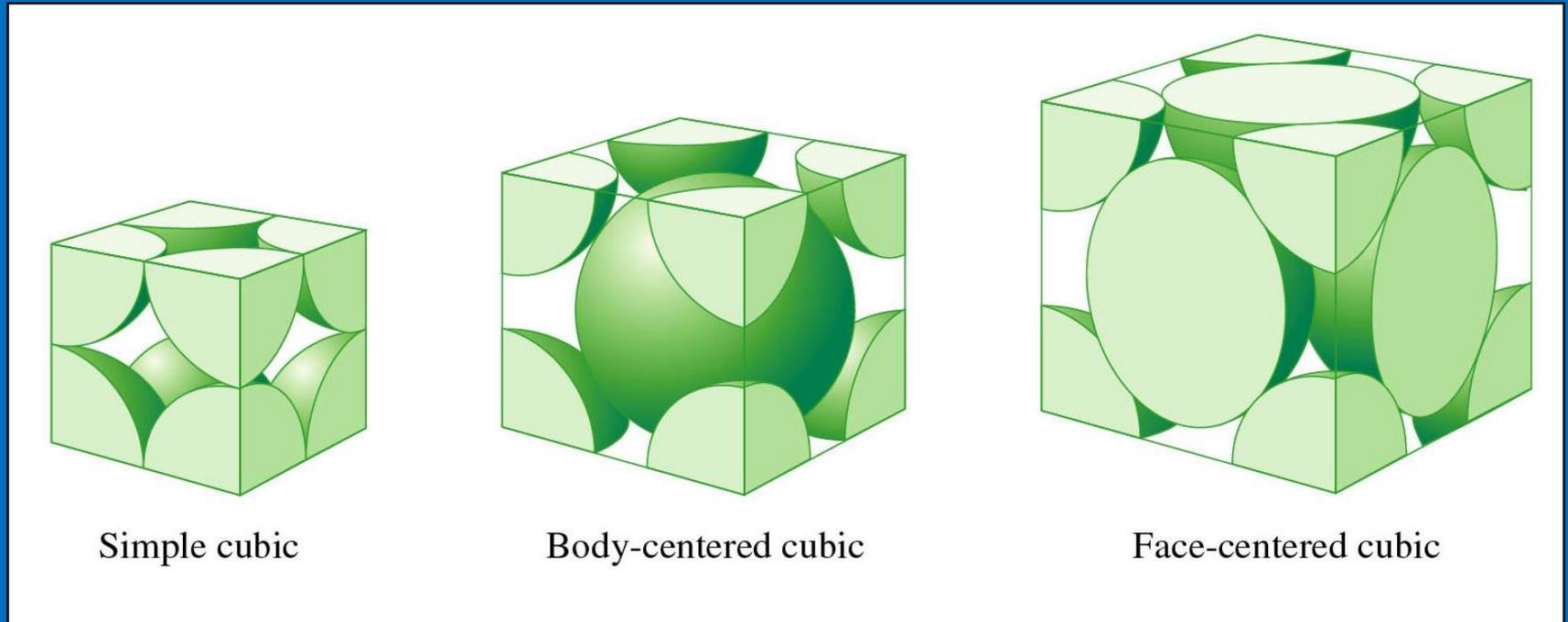
Body-centered cubic



Face-centered cubic

Cubic Unit Cells

- Space-filling representation of cubic unit cells



Simple cubic

Body-centered cubic

Face-centered cubic

1 atom/cell

2 atoms/cell

4 atoms/cell

$$8 \times 1/8 = 1$$

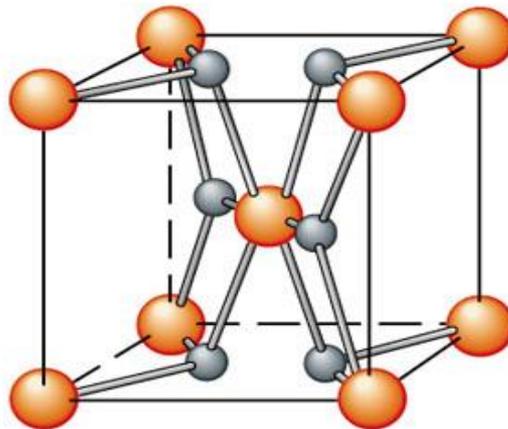
$$8 \times 1/8 + 1 = 2$$

$$8 \times 1/8 + 6 \times 1/2 = 4$$

Practice Problem

Shown here is a representation of a unit cell for a crystal. The orange balls are atom A, and the grey ball are atom B.

- What is the chemical formula of the compound that has this unit cell (A_7B_7)?
- Consider the configuration of the A atoms. Is this a cubic unit cell? If so, which type?



a. A_9B_6

b. **Body Center Cubic**

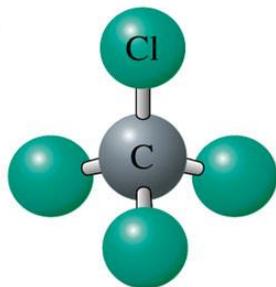
Practice Problem

Decide which substance in each of the following pairs has the lower melting point

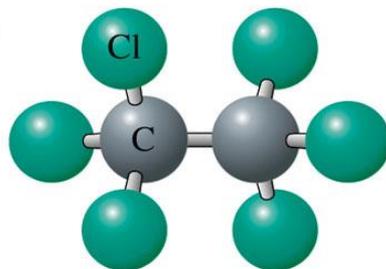
Explain how you made each choice

a. potassium chloride, KCl; or calcium oxide, CaO

b. carbon tetrachloride,



or hexachloroethane,



c. zinc, Zn; or chromium, Cr

d. acetic acid, CH_3COOH ; or ethyl chloride, $\text{C}_2\text{H}_5\text{Cl}$

- KCl (775°C) vs CaO (2572°C), both ionic but calcium oxide is smaller and has a higher charge magnitude requiring more energy to overcome forces holding molecules together
- Carbon tetrachloride (mp -19°C) would have a lower melting point than hexachloroethane (mp-tp 187) because it has a lower molecular weight and less intermolecular forces to overcome
- Chromium (mp 1857°C) has a much high melting point than Zinc (mp 419°C) because of its higher ionic charges and smaller radius
- Acetic acid (mp 16.5°C) has a high mp than Ethyl chloride (mp -139°C) because hydrogen bonds require more energy to overcome the forces holding the molecules together

Crystal Defects

- There are principally two kinds of defects that occur in crystalline substances
 - **Chemical impurities**, such as in rubies, where the crystal is mainly aluminum oxide with an occasional Al^{3+} ion replaced with Cr^{3+} , which gives a red color
 - **Defects in the formation of the lattice.** Crystal planes may be misaligned, or sites in the crystal lattice may remain vacant

Calculations Involving Unit Cell Dimensions

- **X-ray diffraction** is a method for determining the structure and dimensions of a unit cell in a crystalline compound
 - Once the dimensions and structure are known, the volume and mass of a **single atom** in the crystal can be calculated
 - The determination of the mass of a single atom gave us one of the first accurate determinations of **Avogadro's number**

Sample Problem

LiCl crystallizes in a face-centered cubic structure. The unit cell length is 5.14×10^{-8} cm. The chloride ions are touching each other along the face diagonal of the unit cell. The Li^+ ions fit into the holes between the chloride ions.

How many Li^+ ions are there in this unit cell?

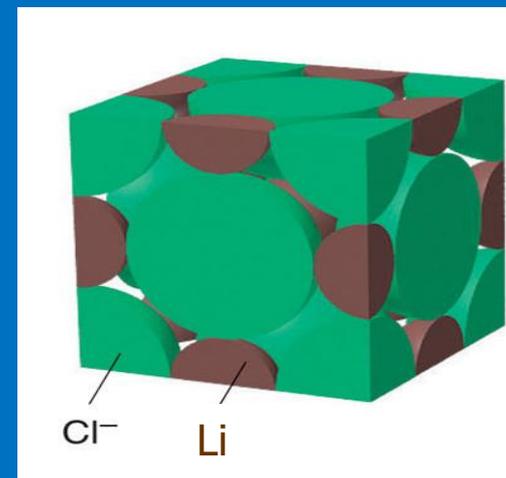
Each cell contains $12 \times \frac{1}{4}$ Li^+ ions
plus 1 whole Li^+ ion

$$12 \times \frac{1}{4} + 1 = 4 \text{ (total)}$$

How many Cl^- ions are there in the cell?

Each cell contain $8 \times \frac{1}{8}$ Cl^- ions and $6 \times \frac{1}{2}$ Cl^- ions

$$8 \times \frac{1}{8} + 6 \times \frac{1}{2} = 4 \text{ (total)}$$



Sample Problem

LiCl crystallizes in a face-centered cubic structure. The unit cell length is 5.14×10^{-8} cm. The chloride ions are touching each other along the face diagonal of the unit cell. The Li^+ ions fit into the holes between the chloride ions.

What is the mass of the LiCl cell?

$$\text{Mass Cl} = 35.45 \frac{\text{g}}{\text{mol}} / (6.02214 \times 10^{23} / \text{mol}) \times 4 \text{ atoms / cell} = 2.355 \times 10^{-22} \text{ g / cell}$$

$$\text{Mass Li} = 6.940 \frac{\text{g}}{\text{mol}} / (6.02214 \times 10^{23} / \text{mol}) \times 4 \text{ atoms / cell} = 4.610 \times 10^{-23} \text{ g / cell}$$

$$\text{Total Mass of Cell} = 2.355 \times 10^{-22} \text{ g} + 4.610 \times 10^{-23} \text{ g} = 2.816 \times 10^{-22} \text{ g}$$

Sample Problem

LiCl crystallizes in a face-centered cubic structure. The unit cell length is 5.14×10^{-8} cm. The chloride ions are touching each other along the face diagonal of the unit cell. The Li^+ ions fit into the holes between the chloride ions.

What is the density of LiCl?

$$\text{Volume of Cell} = (5.14 \times 10^{-8})^3 = 1.36 \times 10^{-22} \text{ cm}^3$$

$$\text{Total Mass of Cell} = 2.816 \times 10^{-22} \text{ g}$$

Density = mass / volume

$$\text{Density} = \frac{2.816 \times 10^{-22} \text{ g}}{1.36 \times 10^{-22} \text{ cm}^3} = 2.07 \text{ g / cm}^3 \quad (\text{Actual Density} = 2.07 \text{ g / cm}^3)$$

Sample Problem

LiCl crystallizes in a face-centered cubic structure

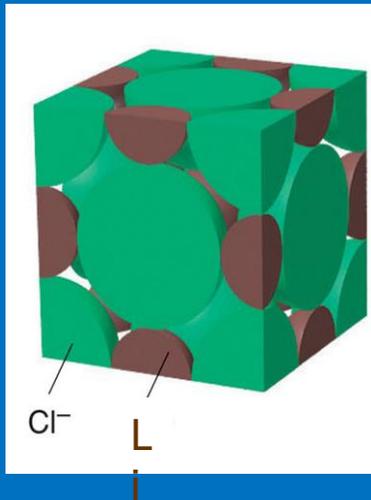
The unit cell length is 5.14×10^{-8} cm

The Chloride ions are touching each other along the face diagonal of the unit cell

The Li^+ ions fit into the holes between the Chloride ions

What is the radius of the Cl^- ion in meters?

(hint: in a square, Diagonal = cell length (S) $\times \sqrt{2}$)



$$\text{Diagonal (C)} = A\sqrt{2} = 5.14 \times 10^{-8} \times 1.4142 = 7.269 \times 10^{-8} \text{ cm}$$

$$C = r + 2r + r = 4r$$

$$r = C/4 = \frac{7.2691 \times 10^{-8} \text{ cm}}{4} \times \frac{1 \text{ m}}{100 \text{ cm}} = 1.8173 \times 10^{-10} \text{ m}$$

Summary Equations

Heat Involved in Temperature Change within a Phase

$$q = n \times C \times \Delta T$$

Heat Involved in Change of Phase

Endothermic

$$q = n \times \Delta H_{\text{vap}} \quad (\text{liquid to gas})$$

$$q = n \times \Delta H_{\text{fus}} \quad (\text{solid to liquid})$$

Exothermic

$$q = n \times (-\Delta H_{\text{vap}}) \quad (\text{gas to liquid})$$

$$q = n \times (-\Delta H_{\text{fus}}) \quad (\text{liquid to solid})$$

Equation Summary

Clausius - Clapeyron Equation

$$\ln P = \left(\frac{-\Delta H_{\text{vap}}}{R} \right) \left(\frac{1}{T} \right) + C$$

$$y = m x + b$$

slope

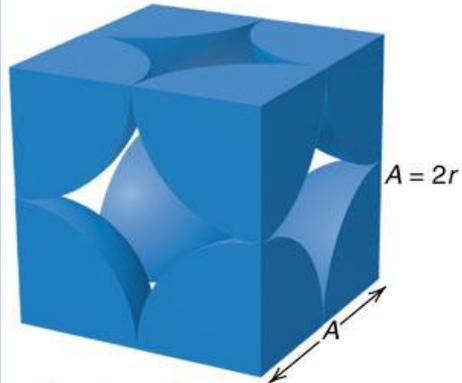
intercept

$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{T_1 - T_2}{T_2 \times T_1} \right)$$

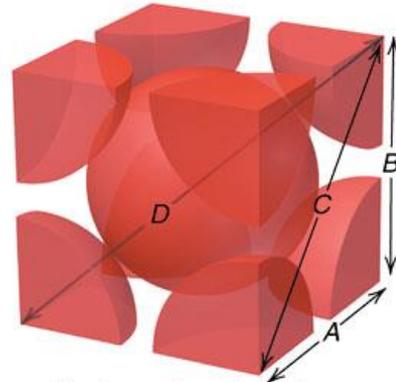
$$R = 8.31447 \text{ J/mol} \cdot \text{K}$$

$$R = 0.0821 \text{ (atm} \cdot \text{L)/(mol} \cdot \text{K)}$$

Equation Summary

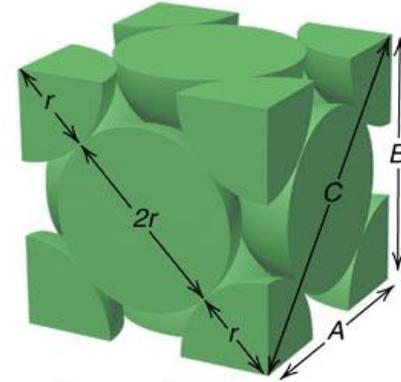


Simple cubic



Body-centered cubic

$$\begin{aligned}C^2 &= A^2 + B^2 = 2A^2 \\D^2 &= A^2 + C^2 = 3A^2 \\D &= \sqrt{3}A = 4r \\A &= \frac{4r}{\sqrt{3}}\end{aligned}$$



Face-centered cubic

$$\begin{aligned}C^2 &= A^2 + B^2 = 2A^2 \\C &= 4r \\16r^2 &= 2A^2 \\A &= \sqrt{8}r\end{aligned}$$