



ScienceGuyz

CHEM 1212 Promo Version

Chapter 10:

Intermolecular Forces: Liquids and Solids

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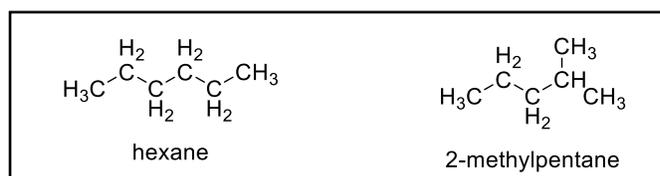
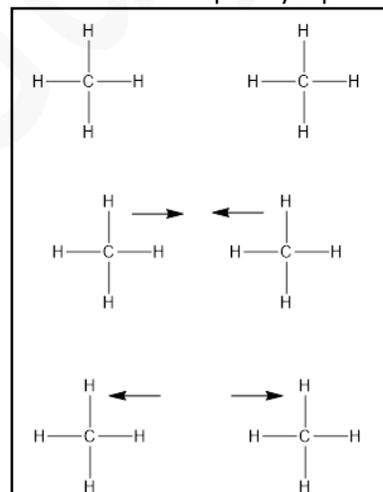
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Learning Objectives: By the end of this workshop, you should be able to:

- Describe what an intermolecular is and why they are important.
- Use principles learned previously (molecular shapes, polarity, and periodic trends) and connect them to intermolecular forces.
- Predict which molecules have the strongest ion-ion interactions when asked to compare.
- Describe polarity and how it arises.
- Describe and recognize intermolecular forces, including dipole-dipole forces, ion-dipole forces, hydrogen bonding, dispersion forces, and induced-dipole-dipole interactions.
- Describe how hydrogen bonding works and what exactly a hydrogen bond is.
- Describe and predict how intermolecular forces impact melting point, boiling point, heat of fusion, heat of vaporization, vapor pressure, surface tension, and viscosity.
- Describe ionic solids and provide examples.
- Describe metallic solids and provide examples.
- Describe covalent solids and provide examples.
- Describe molecular solids and provide examples.
- Describe the different type of cubic unit cells and the different information associated with each.
- Define endothermic and exothermic processes and provide examples of each.
- Define vapor pressure and understand the difference between “boiling point” and “normal boiling point.”
- Predict impacts to vapor pressure by making changes to temperature and pressure/volume.
- Recognize when to use the Clausius-Clapeyron equation and what information it can provide.
- Interpret heating curve diagrams and understand the information that they can provide.
- Define heat of vaporization and heat of fusion.
- Understand when to apply to $q = mc\Delta T$ formula and the information that it provides you.
- Understand when to apply the heat of vaporization and the heat of fusion equations and the information that they provide you.
- Define cohesive and adhesive forces, capillary action, and viscosity.
- Interpret phase diagrams and understand the vocabulary used to describe these diagrams.
- Define and provide examples of dynamic equilibrium.

Intermolecular Forces:

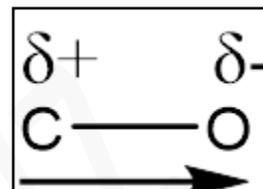
- I. **Intermolecular forces** correspond to the attractive forces *between* individual molecules. Intermolecular forces affect chemistry, as they are related to the physical properties of compounds (melting point, boiling point, enthalpy of vaporization, enthalpy of fusion and solubility). The strength of intermolecular forces of attraction is the fundamental difference between states of matter. **Intramolecular forces** hold atoms together *within* a molecule and include covalent and ionic bonding.
- There are several types of intermolecular forces that we will consider in this workshop, including van der Waals forces (dispersion forces, dipole-dipole interactions, and hydrogen bonding), ion-dipole interactions, ion-induced dipole forces, and dipole-induced dipole forces.
 - All intermolecular forces are weaker than intramolecular forces.
- II. The **dispersion force** (sometimes referred to as **London dispersion force**) is the weakest type of intermolecular force. The dispersion force is a *temporary* attractive force that results when the electrons in two adjacent atoms occupy positions that make the atoms form temporary dipoles.
- Students sometimes struggle with visualizing what exactly dispersion forces are, so let us discuss the image to the right:
 - The image is of the molecule methane (CH_4). We know that methane is a non-polar molecule that has four hydrogen atoms surrounding a central carbon.
 - Ideally, the electrons should be even around the methane molecule, but that is not always the case. For brief moments, there may be an uneven distribution of electrons within the molecule. This creates brief dipoles. If another methane molecule is around (or something else non-polar), the brief dipoles within that molecule can interact with the molecule, shown in the image to the right.
 - NOTE:** ALL substances will experience dispersion forces!
 - Atoms and molecules that are larger will have stronger dispersion forces than smaller atoms and molecules. This is because the larger atoms and molecules have more electrons and are thus more polarizable.
 - Polarizability** is the measure to which the electrons of an atom can be distorted due to another molecule or electric field.
 - Finally, dispersion forces will increase with increasing contact area between molecules. This means that a linear isomer will have greater dispersion forces than a branched isomer.



- III. In **dipole-dipole interactions**, the positive end of one polar molecule and the negative end of another polar molecule are drawn together. Non-polar compounds cannot have dipole-dipole interactions.



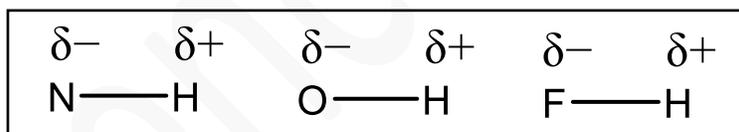
- a. **Remember:** Polar covalent bonds arise when two atoms with large differences in electronegativities are bound together. When two atoms having large differences in electronegativities are bound together, the atoms will share electrons within the bond unevenly.



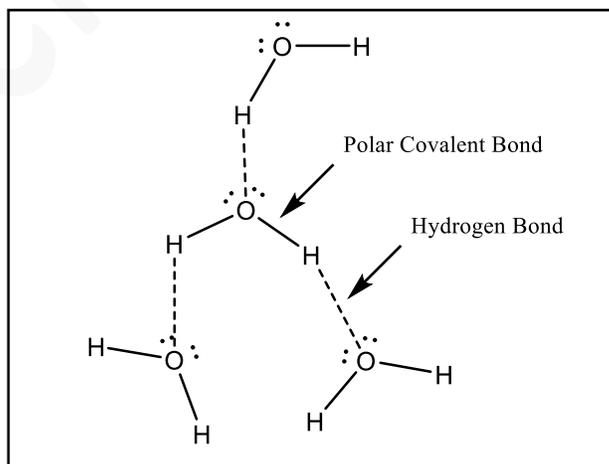
- i. **Partial Negative Charges:** In any polar covalent bond the atom which is more electronegative will bear a partially negative charge (δ^-).
- ii. **Partial Positive Charges:** In any polar covalent bond the atom which is more less electronegative will bear a partially positive charge (δ^+).

- IV. **Hydrogen bonding** occurs when a hydrogen is bound to one of three electronegative elements (oxygen, nitrogen, or fluorine) and partial charges are created on the atoms. The hydrogen possesses the partial positive charge (since it is less electronegative) and the oxygen, nitrogen, or fluorine possesses the partial negative charge (since it is more electronegative).

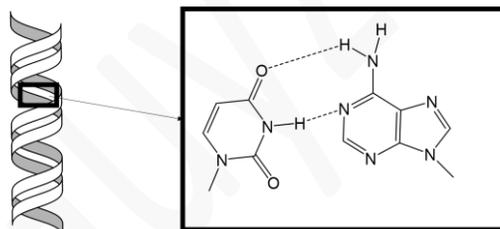
- a. These partial charges will only exist if hydrogen is bound to oxygen, nitrogen, or fluorine. Any molecule which has an X-H bond (where X is O, N, or F) can participate in hydrogen bonding with other molecules that have X-H bonds.



- b. A hydrogen bond is **NOT** the bond between hydrogen and nitrogen, oxygen, or fluorine. This bond is a polar covalent bond. A hydrogen bond is an interaction between partial charges that are a *consequence* of the polar covalent bond between these atoms.



- c. Water's ability to hydrogen bond with other water molecules lead to properties that are essential for life. These properties include:
- Water has a neutral pH and due to its polarity, can effectively dissolve other polar solvents.
 - High heat capacity, which means water is resistant to temperature changes. This is why it takes so long to boil water, even at higher temperatures.
 - More interesting than boiling water is why humans sweat. When you get hot, you begin to sweat. Sweat can transfer heat away via evaporation. This is only effective because sweat is mostly water, and because water can store such a great deal of energy before evaporating. When sweat does acquire enough energy to evaporate, its hydrogen bonds break, and the sweat takes the energy previously contained in those bonds away from the organism.
- d. Ever wondered what holds double-stranded DNA together? This phenomenon is largely due to hydrogen bonding! When DNA helicase unwinds DNA for the DNA replication process, it is breaking the hydrogen bonds that connect the nucleotides together. The base pairs that makeup DNA are held together through hydrogen bonding.



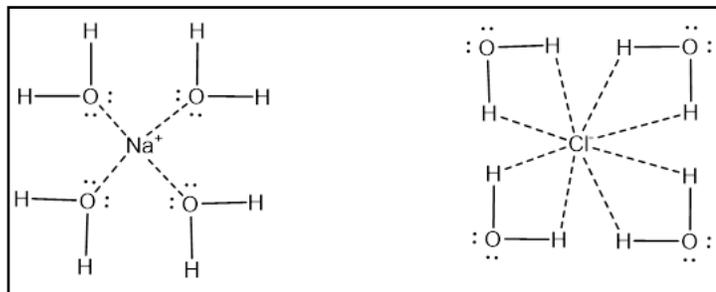
Example: Which of the following molecules can form hydrogen bonds with water?

Molecule	Ability to Hydrogen Bond with Water?
CH_3OCH_3	
HI	
$\text{C}_6\text{H}_5\text{CH}_2\text{F}$	
$\text{C}_6\text{H}_5\text{NH}_2$	
$\text{C}_6\text{H}_5\text{CO}_2\text{H}$	
$\text{N}(\text{CH}_3)_3$	

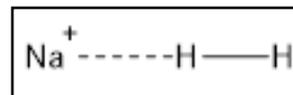
Example: While cooking your favorite pasta (yum), you notice that the water begins to boil. When water boils, _____ interactions are broken and _____ is/are being released.

- Intramolecular; H_2O gas.
- Intramolecular; H_2 and O_2 .
- Intermolecular; H_2O gas.
- Intermolecular; H_2O liquid.
- Intermolecular; H_2 and O_2 .

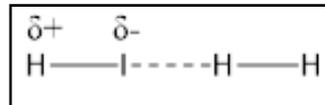
- V. **Ion-dipole forces** result due to the electrostatic interactions between an ion (cation or anion) and the partial charges present in a polar molecule.



VI. Ion-Induced Dipole Forces exist between a nonpolar substance and an ion. The logic for how these forces work is the same as in dispersion forces, however, one of our molecules is an ion in the case instead of them both being nonpolar



VII. Dipole-induced dipole Forces exist between a nonpolar substance and a polar substance. The logic for how these forces work is the same as in dispersion forces, however, one of our molecules is polar in the case instead of them both being nonpolar.



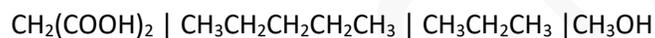
VIII. Summary of Intermolecular Forces: The chart provided below gives a summary of all intermolecular forces and the differences between the.

Rank	Intermolecular Force	Compound Requirement
1	Ion-Dipole Forces	Need ionic compound/ion and a polar molecule. <u>Example:</u> Na ⁺ and H ₂ O
2	Hydrogen Bonding	H must be bound to F, N, or O to participate in hydrogen bonding. <u>Example:</u> H ₂ O, CH ₃ OH
3	Dipole-Dipole Forces	Polar compounds will have dipole-dipole forces. <u>Example:</u> CH ₂ Cl ₂ , CO
4	Ion-Induced Dipole Forces	One ionic compound/ion and one non-polar molecule will induce a dipole. <u>Example:</u> H ₂ and Na ⁺
5	Dipole-Induced Dipole Forces	One polar molecule and one non-polar molecule will induce a dipole. <u>Example:</u> H ₂ and NH ₃
6	London Dispersion Forces	All compounds will have London dispersion forces.

- IX. Relationship Between Intermolecular Forces and Physical Properties of Matter:** Below is a table summarizing the relationship between intermolecular forces and physical properties of matter. We will define each of these physical properties in more detail later in the workshop.

Effect of <i>INCREASING</i> IMFs on Physical Properties of Matter	
Melting Point	↑ Increase
Boiling Point	↑ Increase
Heat of Fusion	↑ Increase
Heat of Vaporization	↑ Increase
Vapor Pressure	↓ Decrease
Surface Tension	↑ Increase
Viscosity	↑ Increase

Example: Ranks the following substances from lowest boiling point to highest boiling point at 1 atm.



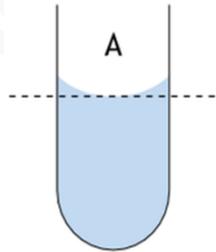
Example: Describe *all* types of intermolecular forces that can occur between particles of each of the following substances.

- XeF₂
- CHCl₃
- He
- CH₃NH₂
- KCl and CO
- O₂ and H₂O

Cohesive and Adhesive Forces and Viscosity:

- I. **Cohesive forces** involve the attraction between adjacent molecules of the same type (such as two water molecules) due to intermolecular forces.
 - a. Cohesive forces are responsible for **surface tension**, a phenomenon where molecules at the surface of water form a tight, lateral barrier and are therefore able to withstand more force without rupturing than would otherwise be expected. The surface tension between molecules at the surface of a liquid must be broken to pierce the liquid. As intermolecular forces increase, surface tension will increase, and vice versa.

- II. **Adhesive forces** involve the attraction between adjacent molecules of a different type (such as the tendency of water to adhere to non-water molecules). This can be stronger than cohesive forces, especially when water is in contact with charged surfaces.

- III. **Capillary action** is the result of adhesive forces. It causes a liquid to “climb” the sides of a tubular space. This happens because a liquid is more attracted to the charged surface of the tube than to other liquid molecules of the same type. For capillary action to occur, adhesive forces between unlike molecules (water and the inside of a glass tube, for example) must be stronger than cohesive forces between like molecules (multiple water molecules, example). This is what creates the “U” of liquids in the glassware and why you measure the amount of liquid in glassware from the center!
 

- IV. **Viscosity** is the resistance of a liquid to flowing. We often think of viscosity as a measure of the *thickness* of a liquid. For example, honey has a higher viscosity than water; it is thicker and therefore flows poorly compared to water. Liquids with higher intermolecular forces will have a higher viscosity, and vice versa.
 - a. The length of a compound also affects viscosity. When comparing two compounds with similar intermolecular forces, the longer compound will have the higher viscosity. This is because longer compounds tend to *tangle* and resist flowing as a result.

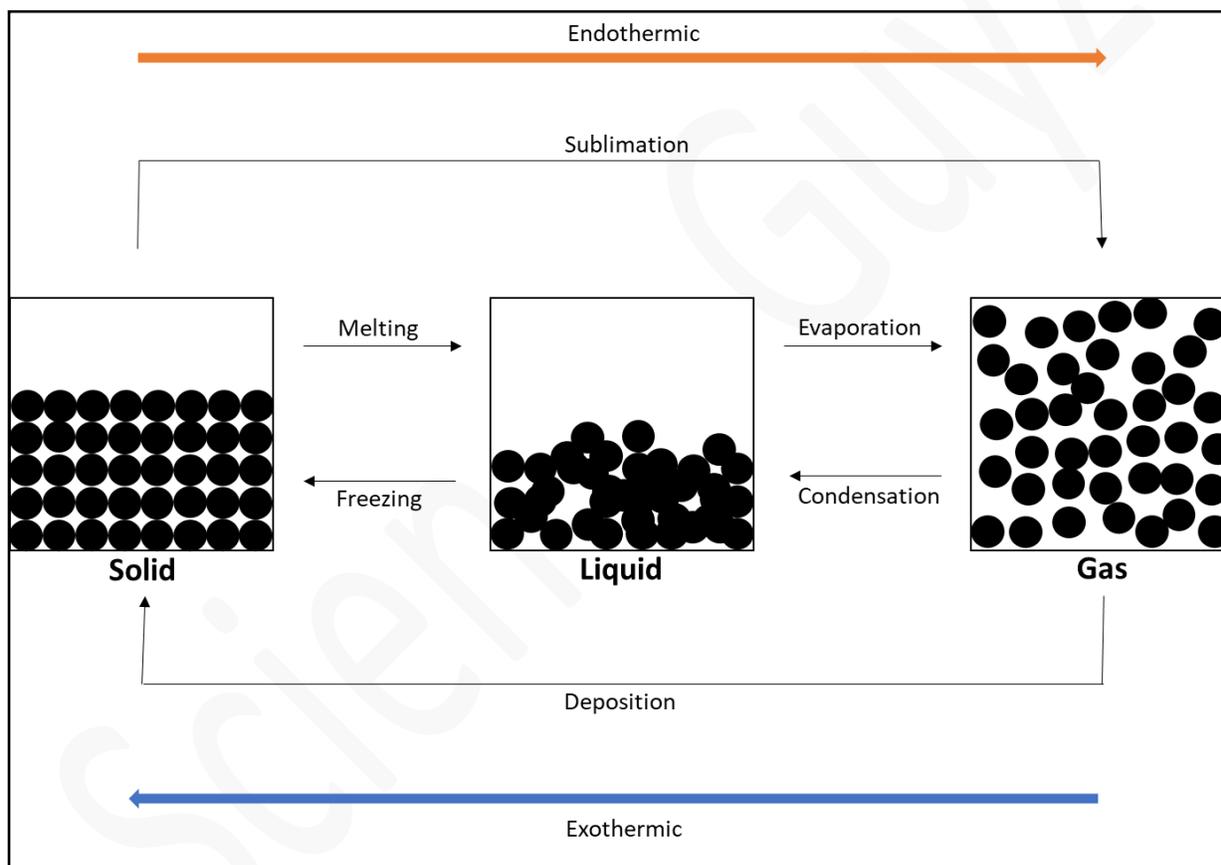
Figure 1: Work by “Jleedev”/Wikimedia Commons
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Example: Which of the follow statements regarding viscosity is/are correct?

- I. Viscosity increases as the temperature decreases.
 - II. Viscosity increases as molecular weight increases.
 - III. Viscosity increases as intermolecular forces decrease.
- a) I only.
 - b) II only.
 - c) III only.
 - d) I and II.
 - e) I, II, and III.

Liquids, State Changes, and Energy:

- I. **Kinetic Energy and Liquids:** The physical properties of liquids are affected by the amount of kinetic energy present in the system. As temperature rises in a sample of liquid (and all matter), the amount of kinetic energy in the sample rises. Before we go any further there is some key vocabulary that we need to review:
- Exothermic Processes:** A process in which a system releases energy in the form of heat to its surroundings. q and ΔH values are negative for exothermic processes.
 - Endothermic Processes:** A process in which a system absorbs energy in the form of heat from its surroundings. q and ΔH values are positive for endothermic processes.
 - As we go from a solid \rightarrow liquid \rightarrow gas, enthalpy and entropy will both increase.



Example: Several phase changes are provided in a chart below. Determine whether each is endothermic or exothermic.

- Melting _____
- Sublimation _____
- Condensation _____
- Vaporization _____
- Freezing _____
- Deposition _____

More on State Changes and Energy:

- V. The energy required to change the temperature and state of a compound depends on the heat capacity of the compound in a particular state and the heat required to effect a change in the state of the compound.
- b. **Heat capacity (C):** The energy released or absorbed by some mass of a substance when the temperature of the substance changes by one degree.
- i. Caution: make sure that the units of C cancel, leaving behind energy in joules or kilojoules. The energy required for the change in temperature of some substance is given by the equation below:

$$Q = mC\Delta T$$

Where m is the mass in grams, C is the specific heat, and ΔT is the change in temperature.

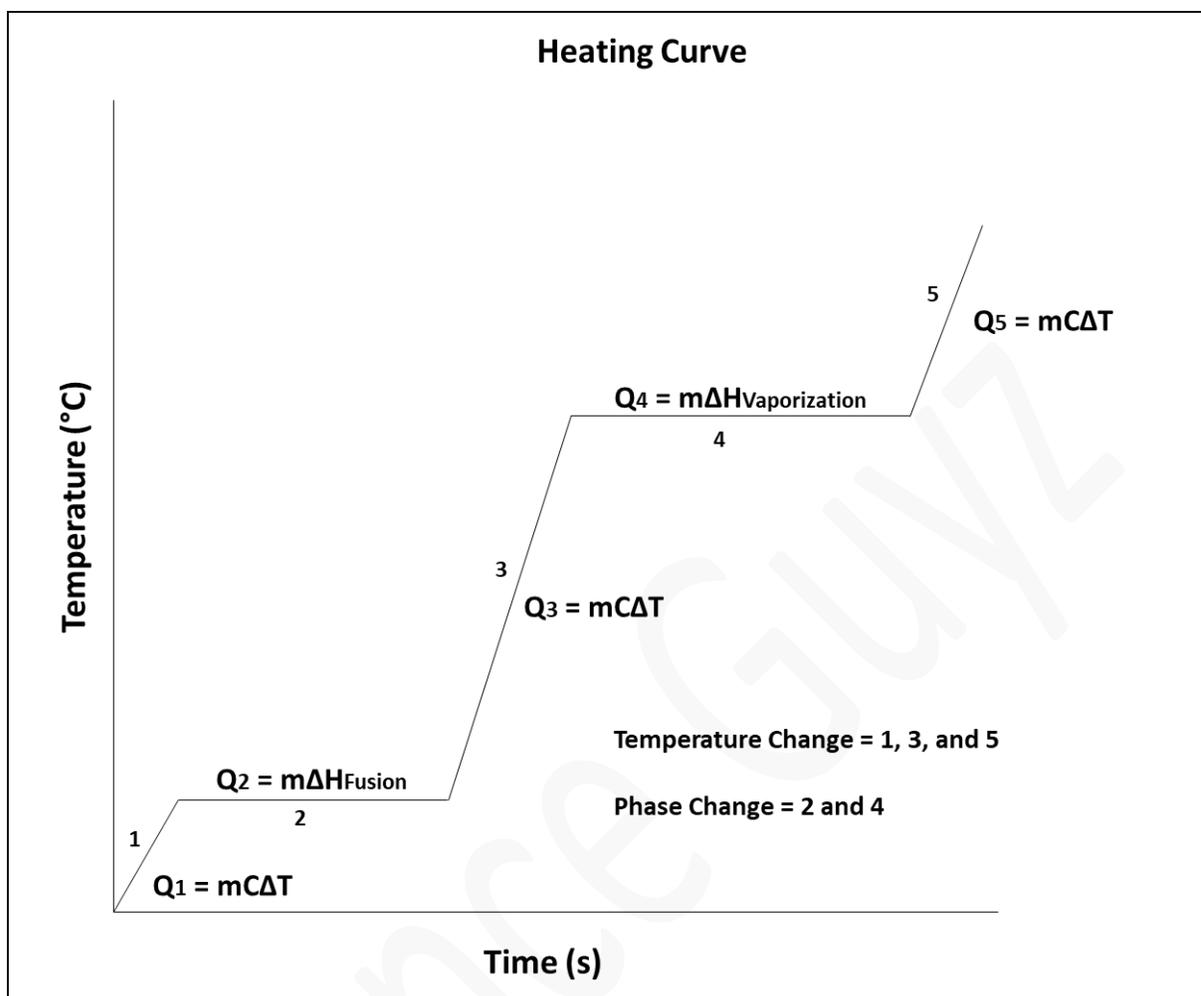
- c. **Heat of Vaporization:** The energy released or absorbed when some mass of compound undergoes a change in state from a liquid to a gas or from a gas to a liquid.
- i. Caution: make sure that the units of ΔH_{vap} cancel, leaving behind energy in joules or kilojoules. Also, pay attention to the sign of your equation!

$$\text{Liquid} \rightarrow \text{Gas: } Q_{\text{state}} = m\Delta H_{\text{vaporization}} \quad \text{Gas} \rightarrow \text{Liquid: } Q_{\text{state}} = -m\Delta H_{\text{vaporization}}$$

- d. **Heat of Fusion:** The energy released or absorbed when some mass of compound undergoes a change in state from a solid to a liquid or from a liquid to a solid.
- i. Caution: make sure that the units of ΔH_{fus} cancel, leaving behind energy in joules or kilojoules. Also, pay attention to the sign of your equation!

$$\text{Solid} \rightarrow \text{Liquid: } Q_{\text{state}} = m\Delta H_{\text{fusion}} \quad \text{Liquid} \rightarrow \text{Solid: } Q_{\text{state}} = -m\Delta H_{\text{fusion}}$$

- e. Note that Heat of Fusion and Heat of Vaporization are not associated with a temperature change since the heat energy is used to increase the distance between molecules.
- f. The total energy for some process is the sum of the energy associated with each process. We use **heating curves** to show the relationship of temperature and heat input (or output) of a system over time. An example of a heating curve and the information that it provides is on the next page.



Practice: How much heat (in Joules) is needed to melt 755 g of ice at 0 °C and then heat the resulting liquid to a vapor at 100 °C. Note: Some potentially useful information is provided in a table below.

Specific Heat of Ice	2.09 J/g°C
Heat of Fusion of ice at 0°C	6.01 kJ/mol
Specific Heat of Liquid Water	4.184 J/g°C
Heat of Vaporization of Liquid Water at 100°C	40.7 kJ/mol
Specific Heat of Steam	2.03 J/g°C

- Using the above heating curve as a reference, we need to determine what all we need to include in our calculations. Since we are starting at 0 °C and adding heat until we get steam at 100 °C, we need to consider points 2, 3, and 4, or $Q_{\text{total}} = Q_2 + Q_3 + Q_4$.

- PAY ATTENTION TO UNITS OF THE INFORMATION GIVEN. It is extremely easy to miss these questions if you do not pay attention to the units of everything provided. Calculations for Q2, Q3, and Q4 are provided.
- Because our ΔH_{Fusion} value is given in kJ/mol, we need to convert our gram amount of ice to mole amount. We can then determine the energy required for this process.

$$\frac{755 \text{ g H}_2\text{O}}{18.02 \text{ g}} \left| \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} \right. = 41.898 \text{ moles}$$

$$Q_2 = m\Delta H_{\text{Fusion}} \rightarrow (41.898 \text{ moles H}_2\text{O}) (6.01 \text{ kJ/mol}) = 251.806 \text{ kJ}$$

The question asks for the final amount of energy to be given in Joules, so we need to convert our Kilojoules to Joules.

$$251.806 \text{ kJ} \rightarrow 251806 \text{ J}$$

- We will now solve for Q3.

$$Q_3 = m\Delta T \rightarrow (755 \text{ g H}_2\text{O})(4.184 \text{ J/g } ^\circ\text{C})(100 \text{ } ^\circ\text{C} - 0 \text{ } ^\circ\text{C}) = 315892 \text{ J}$$

- Because our $\Delta H_{\text{Vaporization}}$ value is given in kJ/mol, we need to convert our gram amount of ice to mole amount. We can then determine the energy required for this process.

$$\frac{755 \text{ g H}_2\text{O}}{18.02 \text{ g}} \left| \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} \right. = 41.898 \text{ moles}$$

$$Q_4 = m\Delta H_{\text{Vaporization}} \rightarrow (41.898 \text{ moles H}_2\text{O}) (40.7 \text{ kJ/mol}) = 1705.244 \text{ kJ}$$

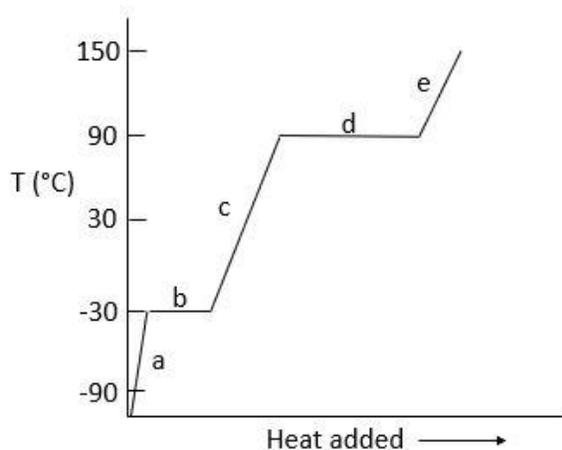
The question asks for the final amount of energy to be given in Joules, so we need to convert our Kilojoules to Joules.

$$1705.244 \text{ kJ} \rightarrow 1705244 \text{ J}$$

- Now we add together Q2, Q3, and Q4 to determine the amount of energy that is required for the described process.

$$Q_{\text{total}} = Q_2 + Q_3 + Q_4 = 251806 \text{ J} + 315892 \text{ J} + 1705244 \text{ J} = \underline{2272942 \text{ J}}$$

Example: Answer the following questions regarding the heating curve of some substance x that is provided below:



1. Which line represents the solid state? The liquid state? The gas state?
2. What is the boiling point of this substance?
3. Which phase has the highest heat capacity for substance x?

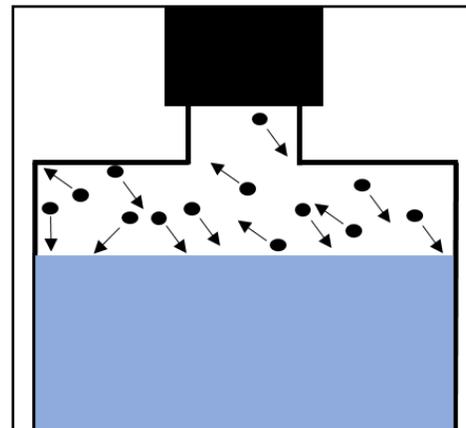
Example: What amount of heat energy (in Joules) must be removed to condense 130.0 g of water vapor at 150 °C to 25 °C?

Specific Heat of Ice	2.09 J/g°C
Heat of Fusion of ice at 0°C	6.01 kJ/mol
Specific Heat of Liquid Water	4.18 J/g°C
Heat of Vaporization of Liquid Water at 100°C	40.7 kJ/mol
Specific Heat of Steam	2.03 J/g°C

Vapor Pressure:

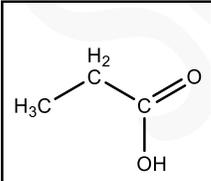
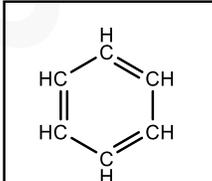
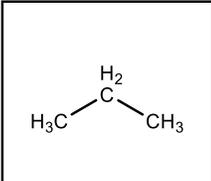
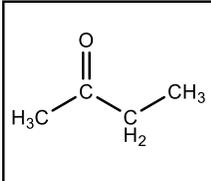
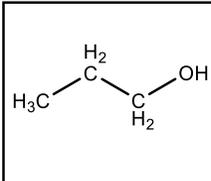
I. **Vapor Pressure** corresponds to the pressure that the gas above a liquid exerts on the liquid when the liquid and gas reach equilibrium, in a closed system. In a closed vessel, the liquid evaporates until the pressure of the gas reaches a certain value. This value is known as the vapor pressure.

- a. As the temperature of a liquid increases the vapor pressure of the liquid will also increase.
 - i. Vapor pressure is dependent on temperature but is independent of both pressure and volume. Adjusting the pressure and volume of a container will not impact the vapor pressure.
- b. The **boiling point** of a liquid is the temperature at which the vapor pressure of a liquid reaches the atmospheric pressure acting on the liquid at a given moment.
 - i. The term **normal boiling point** refers to the temperature at which a will liquid boil when exposed to 1 atm of atmospheric pressure.
- c. The greater the IMFs a compound has the lower the vapor pressure of the compound. The greater the IMFs a compound has the higher the boiling point. Vapor pressure and boiling point are inversely related.

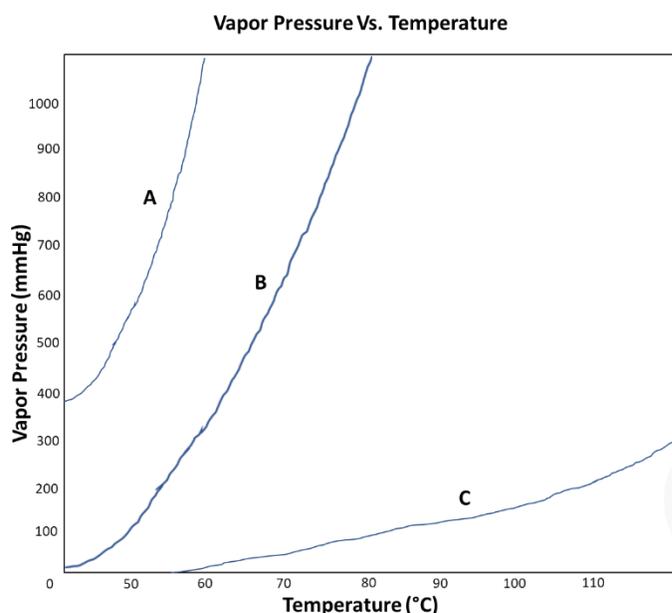


Example: In a general chemistry lab, some 1.0 L flask contains 100 g of some arbitrary liquid. Another 2.0 L flask contains 200 g of the same arbitrary liquid. How do the vapor pressures of the two flasks compare?

Example: Place the following compounds in order of increasing vapor pressure.

 <p>A. Propanoic Acid</p>	 <p>B. Benzene</p>	 <p>C. Propane</p>	 <p>D. 2-Butanone</p>	 <p>E. Propanol</p>	
_____		_____		_____	
Lowest Vapor Pressure				Highest Vapor Pressure	

Example: The graph below shows the relationship between vapor pressure and time for three substances (A, B, and C). Predict the normal boiling point for substance B and rank the three substances from most intermolecular forces to least.



- II. **The Clausius-Clapeyron Equation:** If a plot of vapor pressure at varying temperatures is made for a liquid the plot will fit the Clausius Clapeyron equation.
- Simply put, if you are trying to determine how the vapor pressure changes with changes in temperature and you have the enthalpy of vaporization of the liquid, use this equation.
 - Note if you have a value for all variables except one, you can solve for the unknown variable.

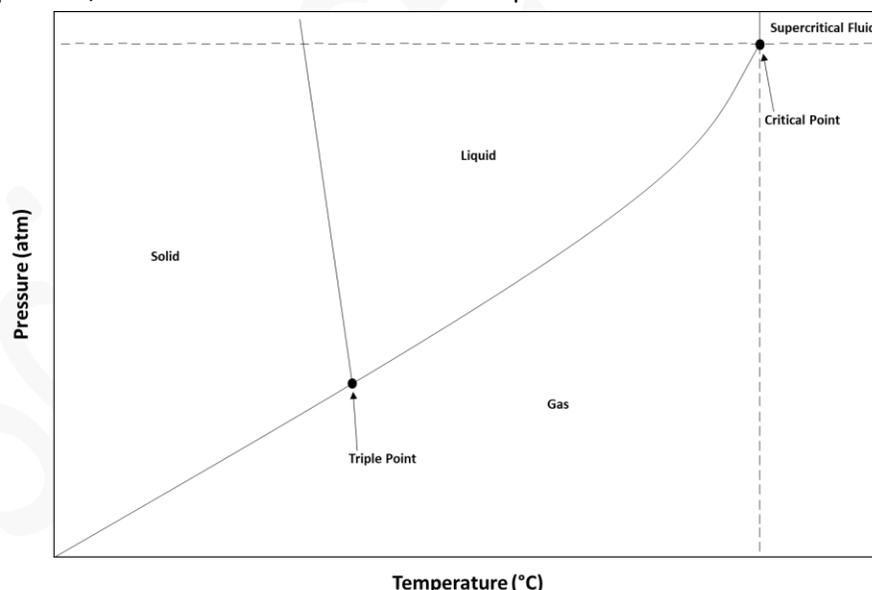
$$\ln \frac{P_2}{P_1} = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

- Where P_1 and P_2 are the pressures in atm, T_1 and T_2 are the temperatures in kelvin, R is the gas constant ($8.314 \times 10^{-3} \frac{\text{kJ}}{\text{K} \cdot \text{mol}}$), and $-\Delta H_{vap}$ is the enthalpy of vaporization in kJ/mol.

Example: At 42 °C, an arbitrary liquid has a vapor pressure of 3.09 torr. Determine the normal boiling point of this arbitrary liquid. Note: The enthalpy of vaporization of the arbitrary liquid is known to be 47.02 kJ/mol.

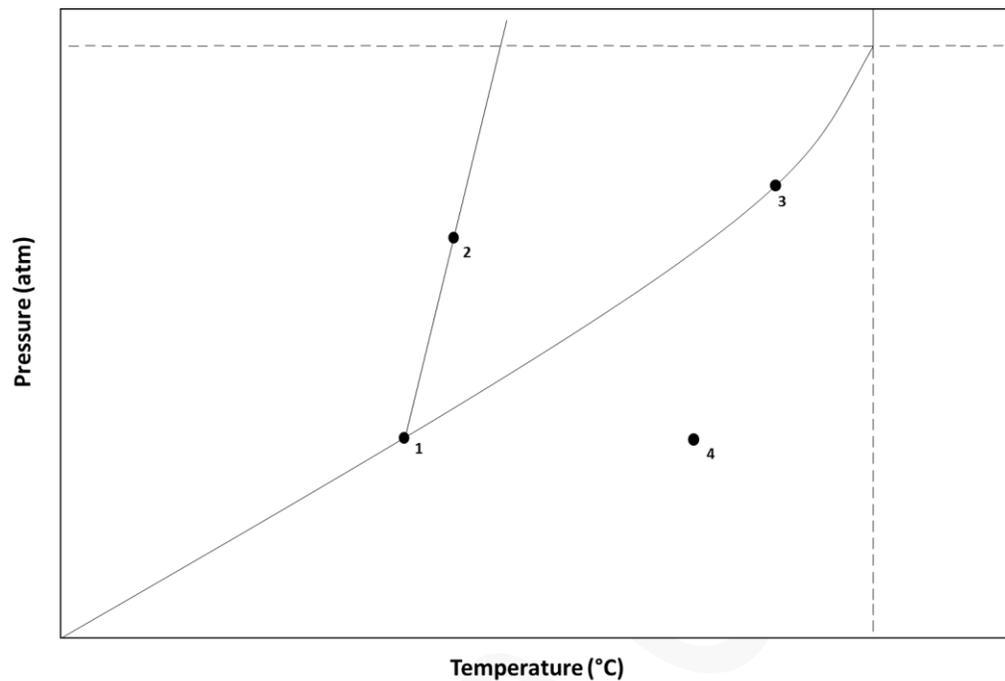
Phase Diagrams:

- I. **Phase diagrams** provide information about the state of a substance at varying temperatures and pressures. The parameters of the phase diagram are shown below:
- Normal Melting Point:** The temperature at which a solid melts at atmospheric pressure (1 atm or 760 mm Hg).
 - Normal Boiling Point:** The temperature at which a liquid boils at atmospheric pressure. (1 atm or 760 mm Hg).
 - Point of Sublimation:** Any temperature and pressure at which a solid is converted directly to a gas, without going through a liquid state.
 - Critical Point:** Occurs under conditions in which no phase boundaries exist. (super critical fluid)
 - Critical Temperature:** The temperature above which vapor cannot be liquefied, no matter how much pressure is applied.
 - Critical Pressure:** The pressure required to liquefy a gas at its critical temperature.
 - Triple Point:** The conditions under which all three phases (solid, liquid and gas) coexist in equilibrium.
 - Density of the Liquid Versus Solid:** To determine whether the liquid of a substance is more or less dense than its corresponding solid, look at the line which acts as the barrier between the solid and liquid states.
 - Negative Slope:** If the slope of the line which partitions the liquid and solid state is negative, the solid will be less dense than the liquid.
 - Positive Slope:** If the slope of the line which partitions the liquid and solid state is positive, the solid will be denser than the liquid.



- II. **Dynamic equilibrium** involves the state of a system where two reciprocal processes are occurring at the same time. An example of these reciprocal processes would be melting and freezing when looking at solids and liquids. When dynamic equilibrium is occurring, there is still a transfer of molecules in the reciprocal processes. The molecules are simply moving from one phase to the other at an equal rate.

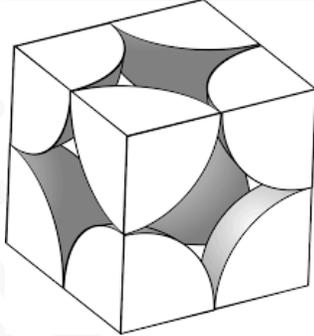
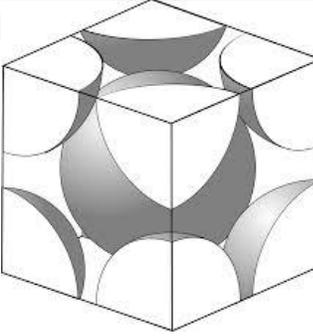
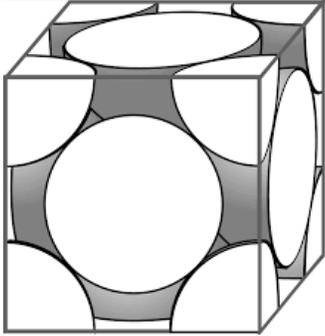
Example: Determine whether the following statements are *TRUE* or *FALSE* regarding the phase diagram shown below. If false, correct the statement to make it true.



1. We would describe point 1 on the phase diagram as the critical point.
2. At point 2 on the phase diagram, the liquid phase is in dynamic equilibrium with the gas phase.
3. At point 3 on the phase diagram, the substance is in the gas phase.
4. At point 4, the substance exists as a gas.
5. The solid form of the substance is said to be denser than the liquid.

An Introduction to Solids:

- I. The two broadest categories of solids are crystalline solids and amorphous solids.
- Crystalline solids** have a highly ordered arrangement of their component atoms, ions, or molecules.
 - The position of the atoms, ions, or molecules that make up crystalline solids is usually described using a **lattice**, which is a three-dimensional system that shows the position of the components of a solid.
 - A **unit cell** is the smallest repeating unit of a lattice. There are several different types of unit cells that you must become familiar with, and these are described below.
 - We will mostly focus on crystalline solids in this workshop.
 - Amorphous solids** have disorder in their structure.
 - A classic example of an amorphous solid that you see every single day is glass.
- II. Solids will crystalize to form three common types of cubic cells: The **simple cubic cell (primitive)**, the **face centered cubic cell**, and the **body centered cubic cell**. Each type of cubic cell will have one or more of the features described below:
- Corners will be composed of 1/8 of an atom.
 - Edges will be composed of 1/4 of an atom.
 - Faces will be composed of 1/2 of an atom.
 - Centers will be composed of 1 atom.
 - The **coordination number** is the number of other “neighbors” or particles that each part of a crystalline solid is in contact with. Simple cubic cells will always have a coordination number of six, body centered cells will always have a coordination number of eight, and face centered will always have a coordination number of twelve.

	Simple Cubic Cell Unit	Body Centered Cubic Unit	Face Centered Cubic Unit
What does this Cell Unit Look Like?			
Number of Atoms in the Cell	1 Atom	2 Atoms	4 Atoms
Radius of the Cell as it Relates to the Length (a)	$a = 2r$	$a = \frac{4r}{\sqrt{3}}$	$a = \frac{4r}{\sqrt{2}}$
Coordination Number	6	8	12

Simple Cubic Unit Cell: Work by "Cdang" and "Samuel Dupre"/Wikimedia Commons
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Body Centered Unit Cell: Work by "Cdang" and "Samuel Dupre"/Wikimedia Commons
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Face Centered Unit Cell: Work by "Cdang"/Wikimedia Commons
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Example: Which unit cell type has the worst packing efficiency? Which has the best packing efficiency?

Example: Determine the radius (in pm) of a body-centered cubic unit cell with an edge length of 600.00 pm?

Example: Some arbitrary metal crystallizes in a simple cubic cell and is observed to have a volume of $7.56 \times 10^{-24} \text{ cm}^3$. Using this information, determine the radius of this arbitrary metal in picometers.

Example: An unknown metal has a face-centered cubic unit cell with the cell edge of 125.5 pm and its molar mass is 107.9 g/mol. Calculate the density of the unknown metal (in g/ml).

III. Characteristics of Types of Crystalline Solids:

- a. Ionic Solids:** These solids are composed of a cation (+) and an anion (-) that are held together by ionic bonds. They will have ions present at their lattices. These solids have high to very high melting points and are hard. They are considered insulators. When dissolved in a liquid they can conduct electricity but do not have this ability in solid form. Examples of ionic solids include MgCl_2 , NaCl , KF .
- Lattice energy** is the energy of formation of a solid crystalline ionic compound from component ions in the gas phase.
 - As a general trend, the higher the charge present on each ion, the larger the lattice energy will be. Decreasing the distance between the atomic nuclei of the ions will increase the strength of the ion-ion interaction. Of the two factors, ion charge is generally more important.
 - Ion-ion interactions influence the boiling point, melting point, heat of fusion, heat of vaporization and vapor pressure.

Effect of Greater Ion-Ion Interactions (Greater Coulomb Force)	
Melting Point	↑ Increase
Boiling Point	↑ Increase
Heat of Fusion	↑ Increase
Heat of Vaporization	↑ Increase
Vapor Pressure	↓ Decrease

Example: Which of the following compounds has the **lowest** melting point?

- LiF
- K_2S
- CaS
- NaCl
- Each of the above compounds has the same melting point.

- b. Molecular Solids:** These solids are composed of neutral atoms that have discrete covalently bonded molecules at their lattices. These solids could be polar or nonpolar, meaning you could see dipole-dipole and/or London dispersion forces present within these solids. These solids have low melting points and mostly soft. They are insulators and do not conduct electricity. Examples include CO_2 , H_2 , S , I_2 , H_2O , C_4H_{10} .
- c. Atomic Solids:** These solids will have atoms at their lattices. Atomic solids are highly variable and thus need to be subdivided into additional categories.
- Metallic Solids:** These solids are formed by metal atoms that are held together by metallic bonds. These solids have variable melting points and variable hardness. Metallic solids malleable, shiny, and ductile and are known to conduct heat and electricity very well. Examples of metallic solids include Cu , Pb , and Ni .
 - Network Solids:** These solids are made up of covalent bonds of elements of electronegative atoms and are often viewed as one giant molecule. These solids have very high melting points and are very hard. Covalent solids do not conduct electricity or heat well and are insulators. Examples include SiC , SiO_2 , and diamond (C).

Characteristics of Types of Crystalline Solids					
Type of Solid	Particles in Unit Cell	IMFs	Melting/Boiling Point	Hardness	Conductivity
Metallic	Metal atoms	Metallic Bonds	Wide range	Variable; malleable and ductile	Good
Ionic	Cations and Anions	Ion-ion (electrostatic)	High to very high	Hard	Nonconducting solids and conducting liquids
Molecular	Molecules	Dispersion, Dipole-Dipole and H-Bonding	Low	Variable, but mostly soft	Nonconducting
Network Covalent	Network of atoms	Covalent Bonds	Very high	Very hard	Nonconducting

Example: Below are a list of different solids and their proposed classifications. Identify the proposed classifications that are correct.

- I. MgCl_2 (s) is an ionic solid.
- II. U (s) is a metallic solid.
- III. I_2 (s) is a molecular solid.
- IV. SiO_2 (s) is a network solid.

- a) Only I is a correct classification.
- b) I and IV are correct classifications.
- c) I, III, and IV are correct classifications.
- d) I, II, and IV are correct classifications.
- e) I, II, III, and IV are all correct classifications.

