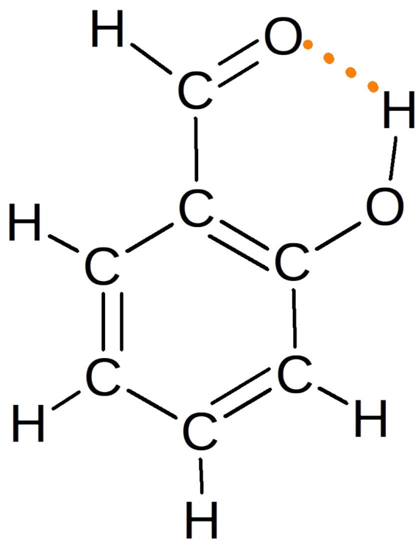
Section 10.1
Intermolecular Forces

Michael Stogsdill

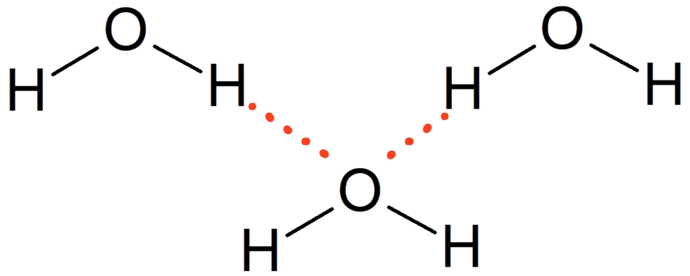
Mott Community College

Intramolecular vs. **Inter**molecular

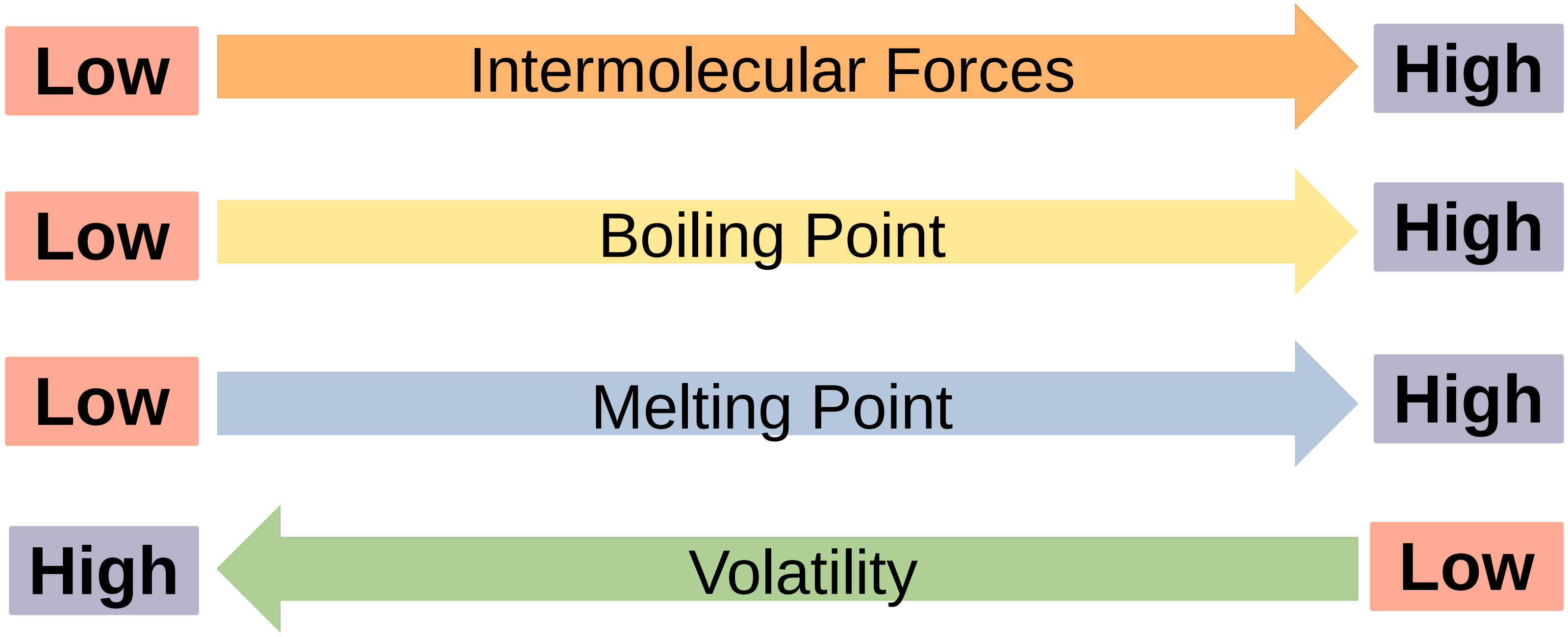
- *Intramolecular Forces* are interactions that occur within a molecule.



- *Intermolecular Forces* are interactions that occur between two or more molecules.



Effect on Bulk Properties



PHET States of Matter

Types of Intermolecular Forces

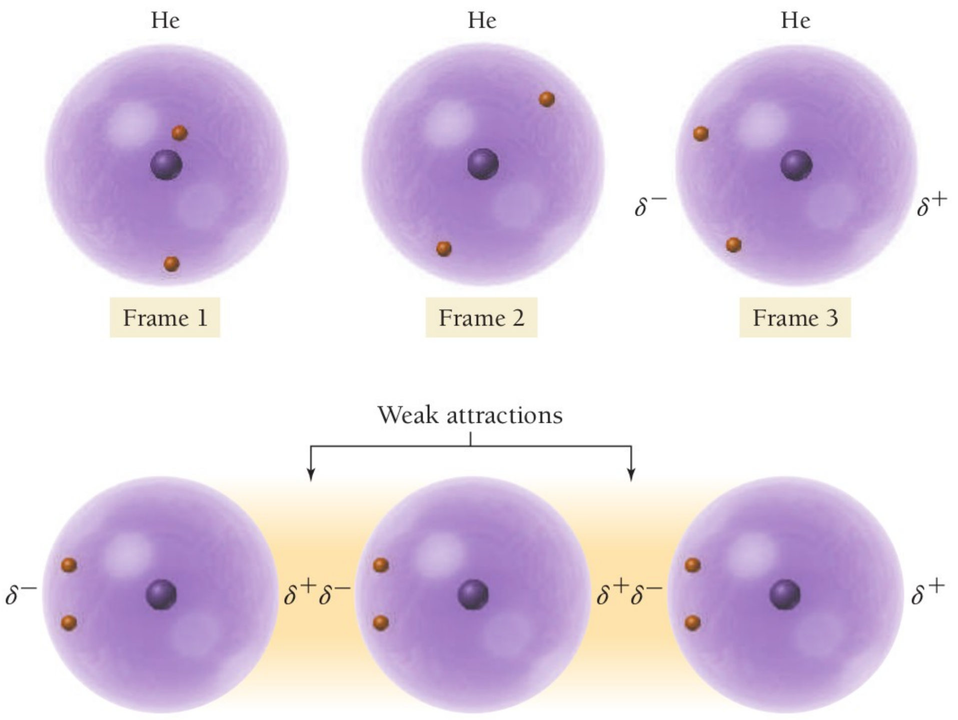
Interaction	Cause
Dispersion (Van Der Waal)	Instantaneous Dipole
Dipole-Dipole	Dipole Moments
Hydrogen Bonding	H interacting with F, O, or N.

Increasing Strength

Dispersion Forces

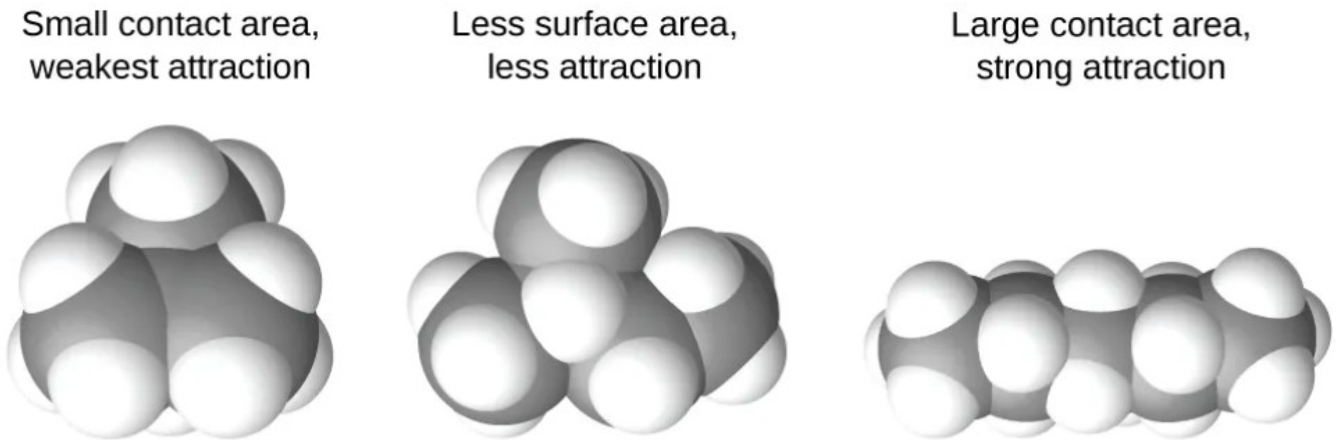
Instantaneous Dipoles

- Instantaneous Dipoles can be induced in any molecule or atom.
- This a weak and non-permanent interaction.
- The interaction becomes stronger with increased formula mass, atomic mass, and effective surface area.



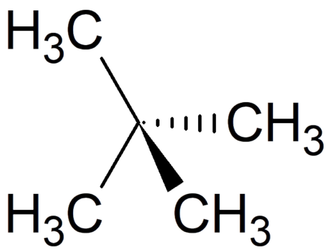
Nonpolar Organic Molecules

- Organic Molecules that only experience dispersive intermolecular interactions.
- Larger molecules have stronger interactions.
- Branching decreases dispersive interactions.



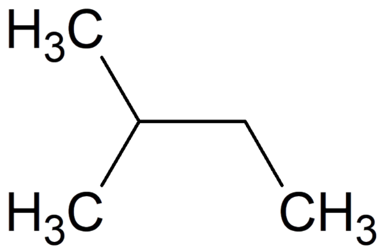
Experimental Evidence

- As predicted, the boiling point increases for less branched constitutional isomers.



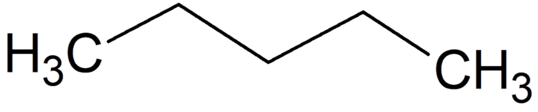
2,2-dimethylpropane

BP : 9.5 °C



2-methylbutane

BP : 27 °C



n-pentane

BP : 36 °C

Dipole Interactions

Periodic Table of Electronegativity

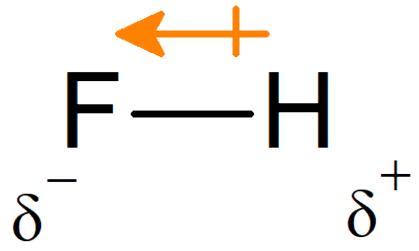
Increasing electronegativity →

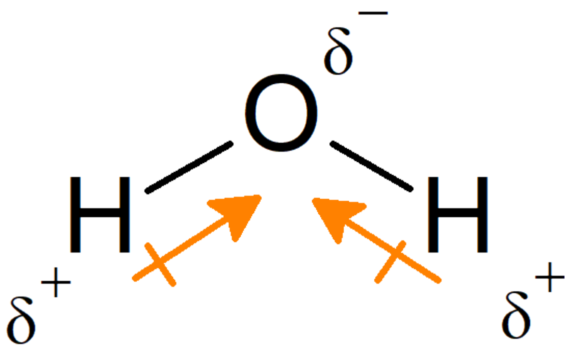
Decreasing electronegativity ↓

			H 2.1														
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										

Bond Dipoles

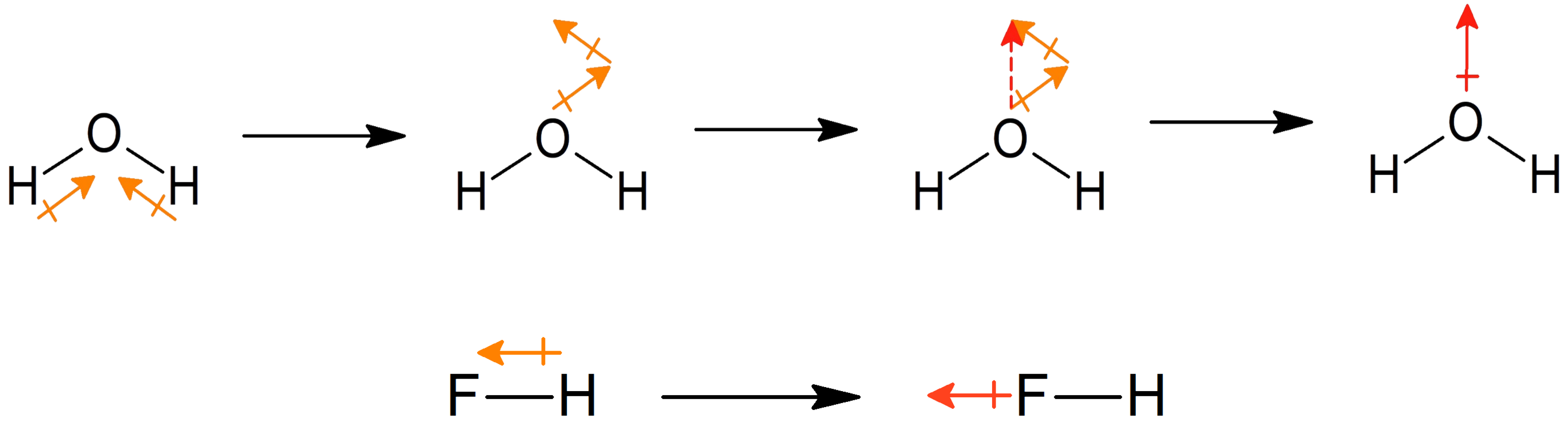
- **Bond Dipoles** are an *intramolecular* interaction that lead to **Dipole Moments** an *intermolecular* interaction.
- They arise when two atoms with different electronegativity are bound together.





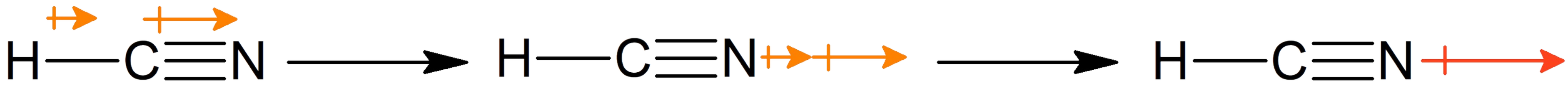
Dipole Moments

- **Dipole Moments** are found by summing a molecule's **Bond Dipole** vectors.



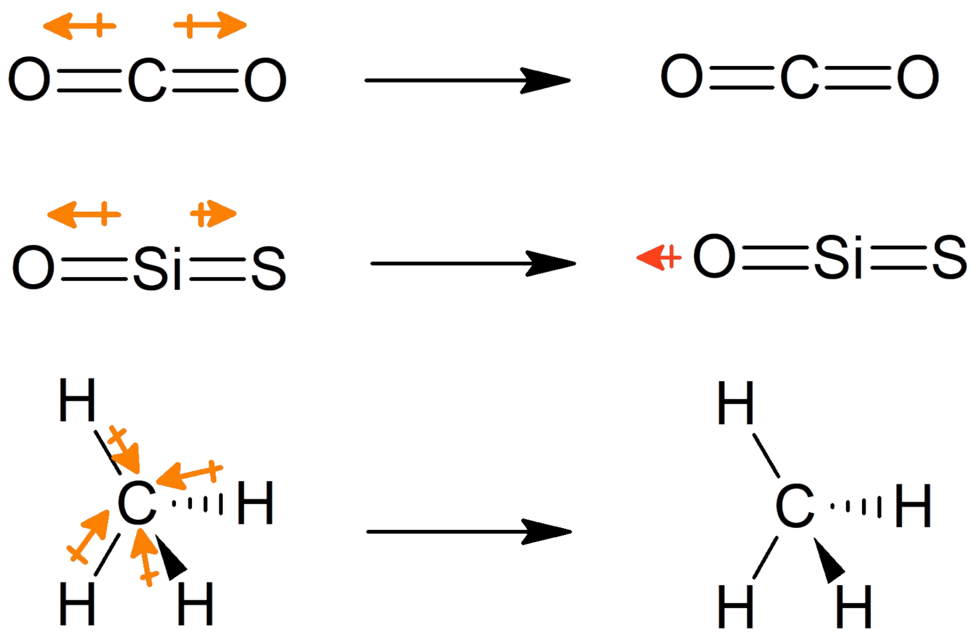
Additive Vectors

- When Bond Dipoles are facing the same direction, the result is a Dipole Moment with a large magnitude.



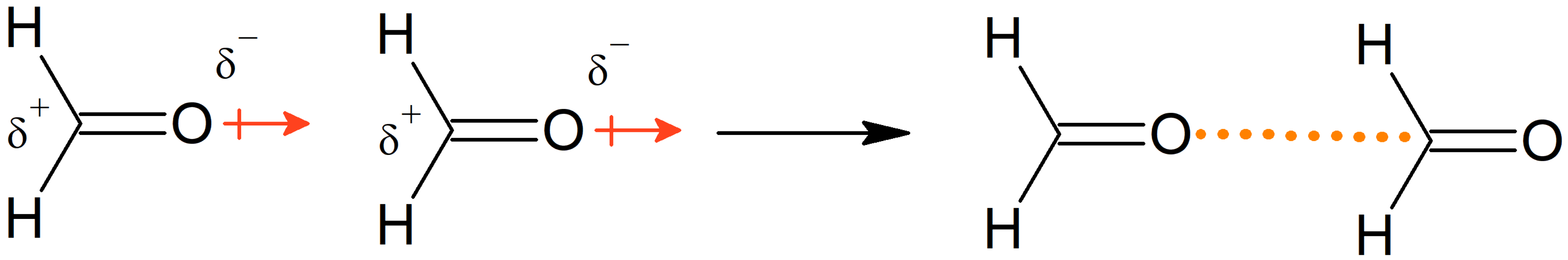
Subtracting Vectors

- When **Bond Dipoles** oppose one another the result can be a small or nonexistent **Dipole Moment**.



Polar Organic Molecules

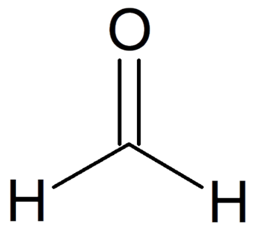
- When two molecules with dipoles come close to one another, their dipoles align leading to an interaction.



PHET Molecule Polarity

Evidence of Dipole-Dipole Interactions

- Again, we can find examples that reinforce our theory.



formaldehyde

Formula Mass:

30.031 g/mol

BP: -19 °C

$H_3C{-}CH_3$

ethane

Formula Mass:

30.07 g/mol

BP: -89 °C

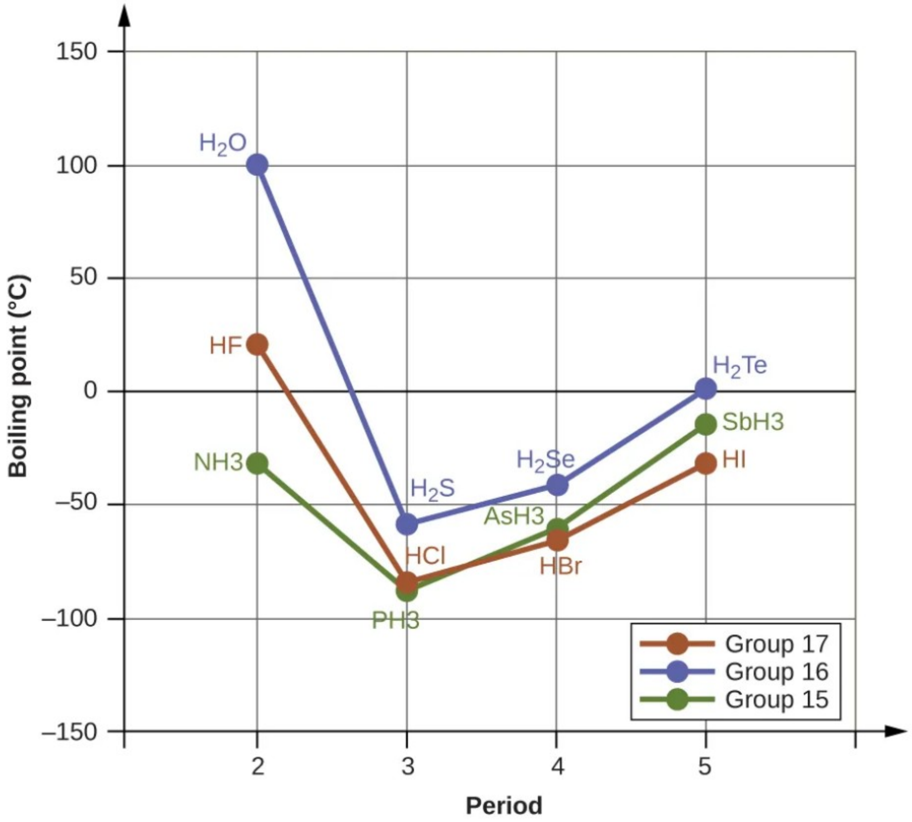
Hydrogen Bonding

Hydrogen Bonding

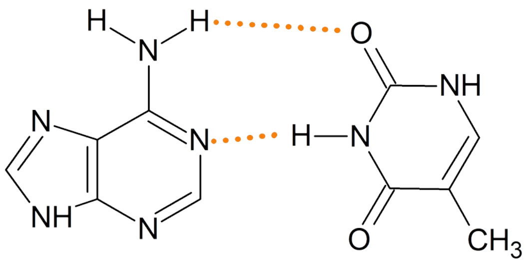
- Hydrogen Bonding is a special case of Dipole-Dipole interactions.
- When hydrogen is bound to a sufficiently electronegative atom (O, N, or F) the single electron on the atom is drawn far away from the nucleus.
- The result is *highly concentrated partial charge* leading to a stronger than normal interaction.

Evidence of Hydrogen Bonding

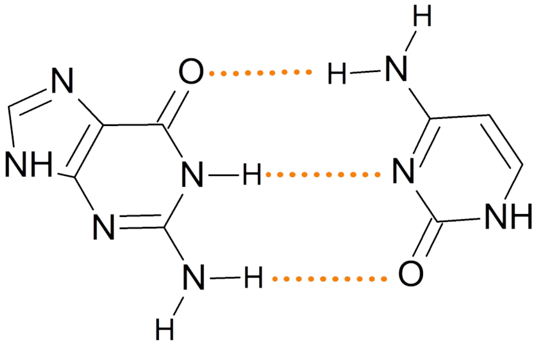
- When we analyze the boiling points of hydrides in Groups 17, 16, and 15, we see a significant break in the trend predicted by considering dispersive forces alone.



Implications in Organic Chemistry



Adenine Thymine



Guanine Cytosine

- Hydrogen Bonding plays an important role in dictating the conformation of many large organic molecules.
- It is responsible for the bonding between nucleic acid base pairs that bind together strands of DNA.

Conclusions

- All molecules and atoms are subject to dispersive forces that induce partial charge differentials leading to *instantaneous dipoles.*
- Differences in the electronegativity of atoms bound together lead to *bond dipoles*.
- Bond dipoles can sum to produce *dipole moments* depending on molecular geometry.
- Hydrogen bound to O, N, or F results in *hydrogen bonding.*
- Intermolecular forces have a significant impact on bulk properties.

Section 10.2
Properties of Liquids

Michael Stogsdill

Mott Community College

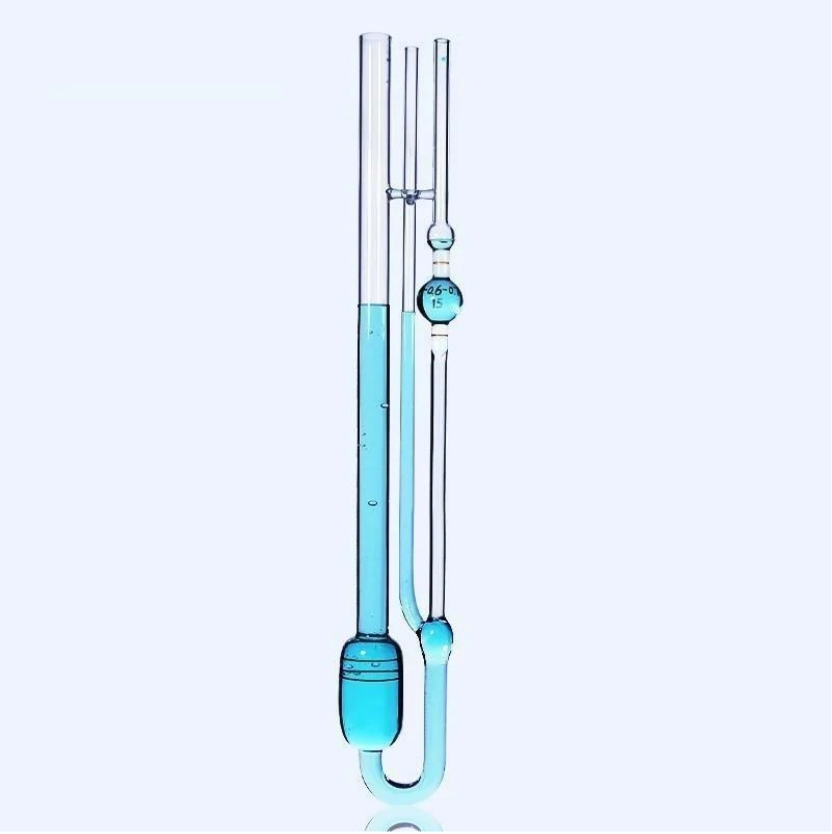
Learning Objectives

- Distinguish between adhesive and cohesive forces
- Define viscosity, surface tension, and capillary rise
- Describe the roles of intermolecular attractive forces in each of these properties/phenomena

Viscosity

- Viscosity is a measure of a liquid’s resistance to flow.
- Low viscosity liquids include: water, gasoline, or acetone
- High viscosity liquids include: honey, motor oil
- It is measured in a number of ways
  - Rotating a spindle in a liquid and measuring the force required to maintain a certain rpm
  - measuring the amount of time it takes for the liquid to pass through a thin tube.

Viscometers



Viscosities of Common Substances

Substance	Formula	Viscosity (mPa·s)
water	H_2O	0.890
mercury	Hg	1.526
ethanol	C_2H_5OH	1.074
octane	C_8H_{18}	0.508
ethylene glycol	$CH_2(OH)CH_2(OH)$	16.1
honey	variable	~2,000–10,000
motor oil	variable	~50–500

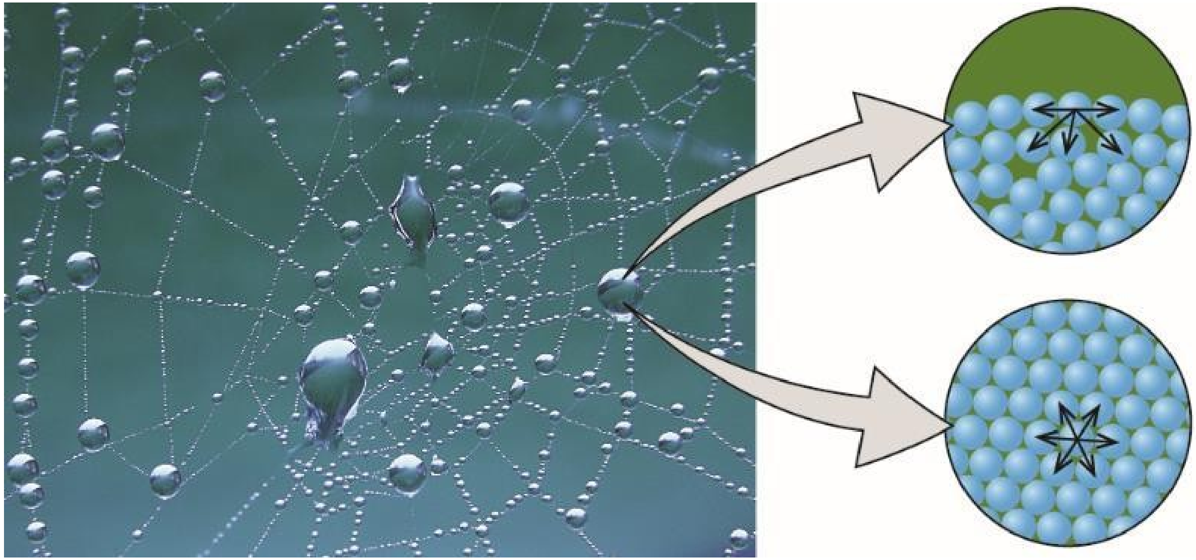
Table 10.2

Cohesives Forces vs. Adhesive Forces

- **Cohesives Forces** are attractive forces between *identical molecules.*
- **Adhesive Forces** are attractive forces between *different molecules.*
- Both of these types of forces are the result of intermolecular forces.

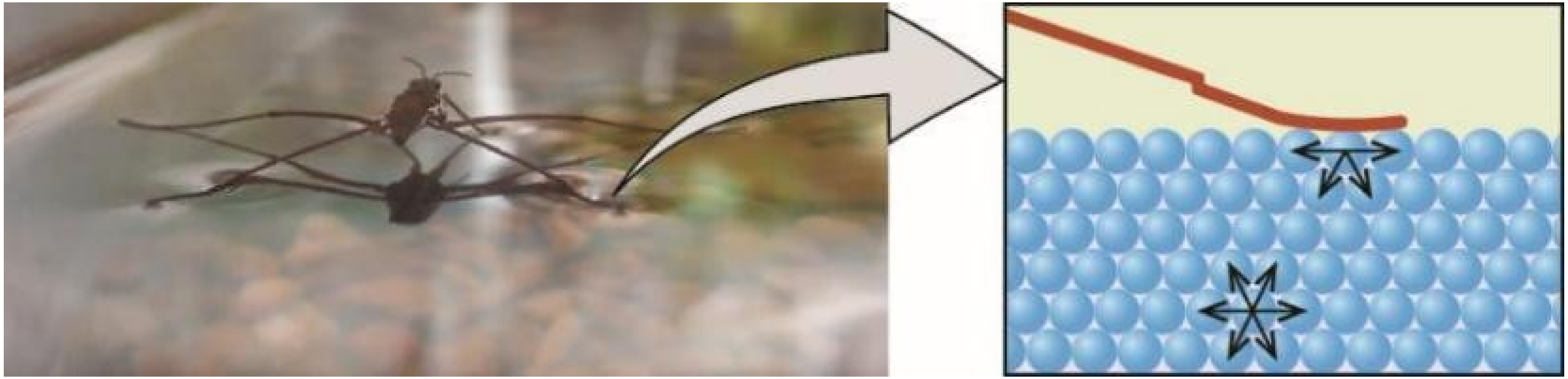
Surface Tension

- **Surface tension** is defined as the energy required to increase the surface area of a liquid, or the force required to increase the length of a liquid surface by a given amount.



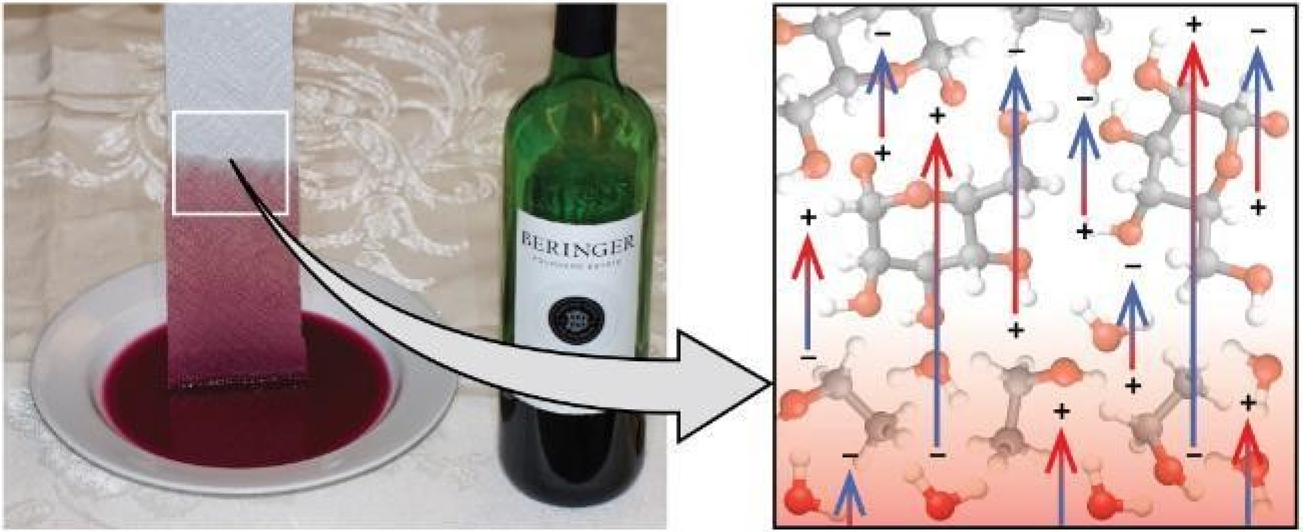
Surface Area Continued

- Surface tension results in a sort of "skin" on the outside of a liquid as it requires force to distort the geometry of the liquids surface.



Capillary Action

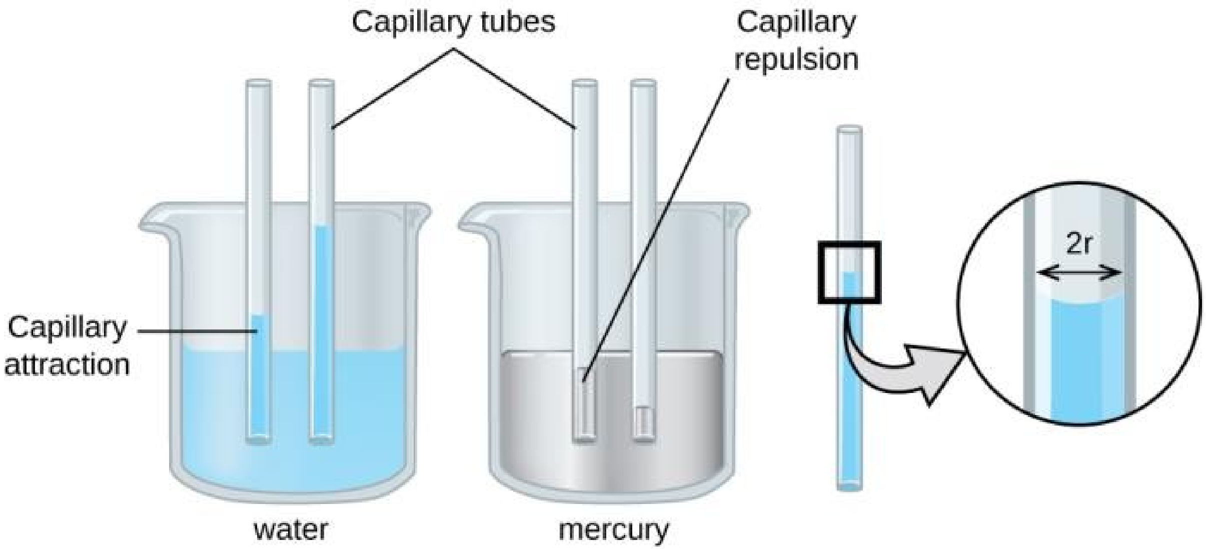
- Adhesive forces cause a phenomenon called capillary action.
- Attractive forces between a liquid and surface will cause a liquid to “climb” up a container wall or “wick” up into a material.



Calculating Capillary Rise

- We can calculate the degree of capillary rise. There are a few more variables here than you may be used to.

$$h = \frac{2T\cos\theta}{r\rho g}$$



Conclusions

- Cohesive forces will cause a liquid to attract itself causing droplets, surface tension, and negative capillary rise.
- Adhesive forces will cause a liquid to be attracted to surfaces causing positive capillary rise and wetting of surfaces.
- Both cohesive and adhesive forces are caused by intermolelcular forces.

Section 10.3
Phase Transition

Michael Stogsdill

Mott Community College

Learning Objectives

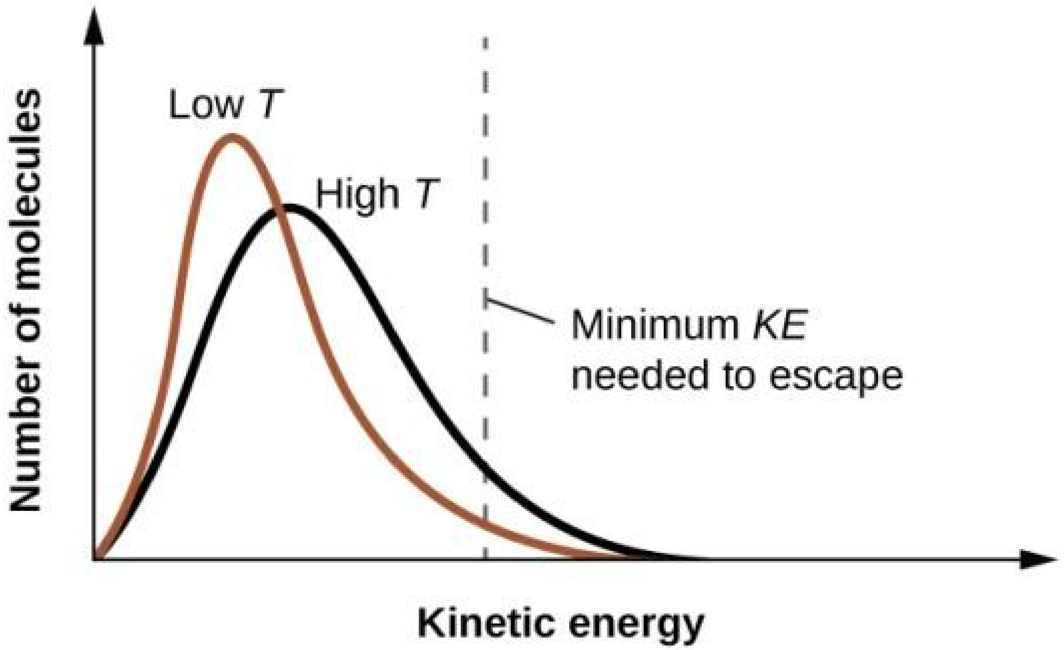
- Define phase transitions and phase transition temperatures
- Explain the relation between phase transition temperatures and intermolecular attractive forces
- Describe the processes represented by typical heating and cooling curves, and compute heat flows and enthalpy changes accompanying these processes

Condensation and Vaporization

- **Vaporization** is the phase transition from a liquid to a gas.
- **Condensation** is the phase transition from a gas to a liquid.
- This process is a **Dynamic Equilibrium**. Molecules are always moving between the two phases.
- The **Vapor Pressure** will determine where this equilibrium lies. It depends on temperature.
- When the **Partial Pressure** of a compound is equal to the **Vapor Pressure**, equilibrium is reached.

Vapor Pressure and Temperature

- Temperature affects the distribution of kinetic energies for the molecules in a liquid. At the higher temperature, more molecules have the necessary kinetic energy, KE, to escape from the liquid into the gas phase.

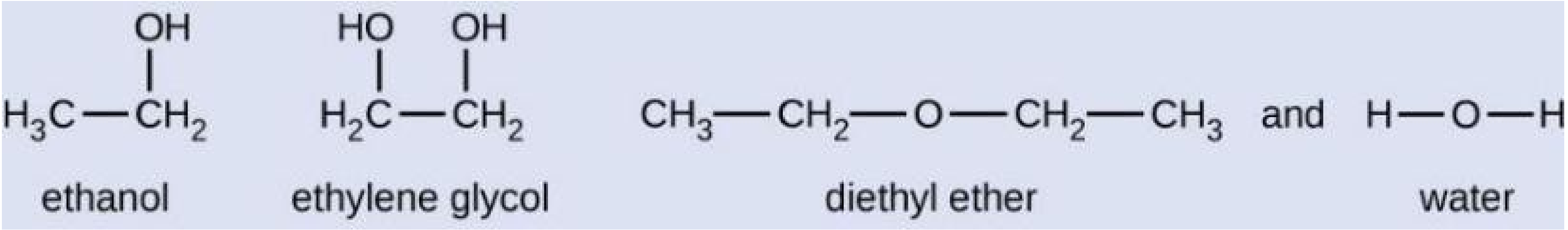


Boiling

- **Boiling** occurs when the vapor pressure of a liquid exceeds the atmospheric pressure above the liquid.
- **Boiling Point** is the temperature at which this occurs. It depends on the atmospheric pressure.
- **Normal Boiling Point** is the boiling point at 1 atm of pressure.
- During boiling the temperature remains constant.

Intermolecular Forces at Work

- When the intermolecular forces between a liquid’s molecules are stronger than the intermolecular forces between the liquid molecules and air molecules, the **vapor pressure** will be relatively low.
- When the intermolecular forces between the air and liquid are stronger than the IMFs between liquid molecules, the vapor pressure will be relatively high.



Clausius-Clapeyron Equation

$$\ln P = -\frac{\Delta H_{\text{vap}}}{RT} + \ln A$$

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

$$R = 8.314\ \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

Enthalpy of Vaporization

- Remember enthalpy from thermochemistry? ΔH_{vap} is a special case of a phase change from liquid to gas.
- Because gases are higher energy species than liquids the process is endothermic.
- The sign of ΔH_{vap} indicates the direction.

$$H_2O(l) \longrightarrow H_2O(g) \qquad \Delta H_{\mathrm{vap}} = 44.01\ \mathrm{kJ/mol}$$

$$H_2O(g) \longrightarrow H_2O(l) \qquad \Delta H_{\mathrm{con}} = -\Delta H_{\mathrm{vap}} = -44.01\ \mathrm{kJ/mol}$$

Melting and Freezing

- **Melting** describes the phase transition from a solid to a liquid.
- **Freezing** describes the phase transition from a liquid to a solid.
- The temperature at which this transition happens is called the **Melting Point** or **Freezing Point.**
- ΔH_{fus} data can be found in Appendix G of the Textbook.

$$H_2O(s) \longrightarrow H_2O(l) \qquad \Delta H_{\text{fus}} = 6.01\ \text{kJ/mol}$$

$$H_2O(l) \longrightarrow H_2O(s) \qquad \Delta H_{\text{frz}} = -\Delta H_{\text{fus}} = -6.01\ \text{kJ/mol}$$

Sublimation and Deposition

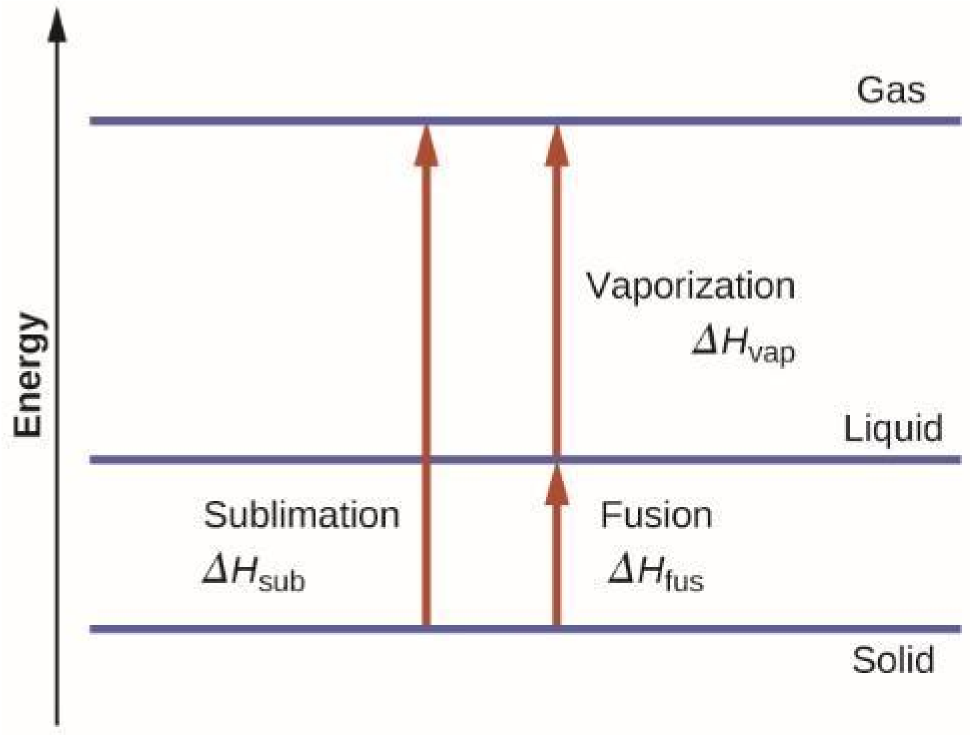
- **Sublimation** describes the transition from a solid to a gas.
- **Deposition** describes the transition from a gas to a solid.
- Whether or not a compound will melt or sublime depends on the its identity, the pressure, and the temperature.

$$\mathrm{CO_2}(s) \longrightarrow \mathrm{CO_2}(g) \qquad \Delta H_{\mathrm{sub}} = 26.1\ \mathrm{kJ/mol}$$

$$\mathrm{CO_2}(g) \longrightarrow \mathrm{CO_2}(s) \qquad \Delta H_{\mathrm{dep}} = -\Delta H_{\mathrm{sub}} = -26.1\ \mathrm{kJ/mol}$$

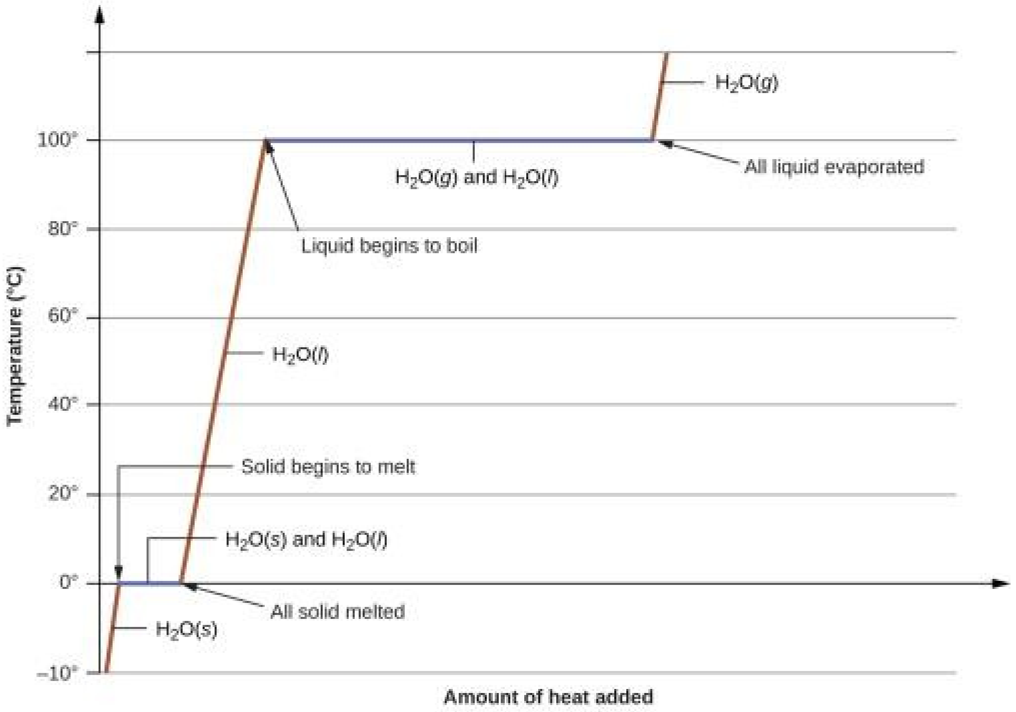
Phase Transition Enthalpy

- For a given substance, the sum of its enthalpy of fusion and enthalpy of vaporization is approximately equal to its enthalpy of sublimation.



Heating Curves

- A typical heating curve for a substance depicts changes in temperature that result as the substance absorbs increasing amounts of heat. Plateaus in the curve (regions of constant temperature) are exhibited when the substance undergoes phase transitions.



Conclusions

- Each type of phase transition has its own enthalpy value associated with it.
- The sign of these enthalpy changes indicate direction. The magnitude indicates the strength of IMFs.
- The Clausius-Clapeyron equation relates the P, T, and ΔH_{vap}.
- The temperature of a compound will remain constant during phase transitions.

Section 10.4
Phase Diagrams

Michael Stogsdill

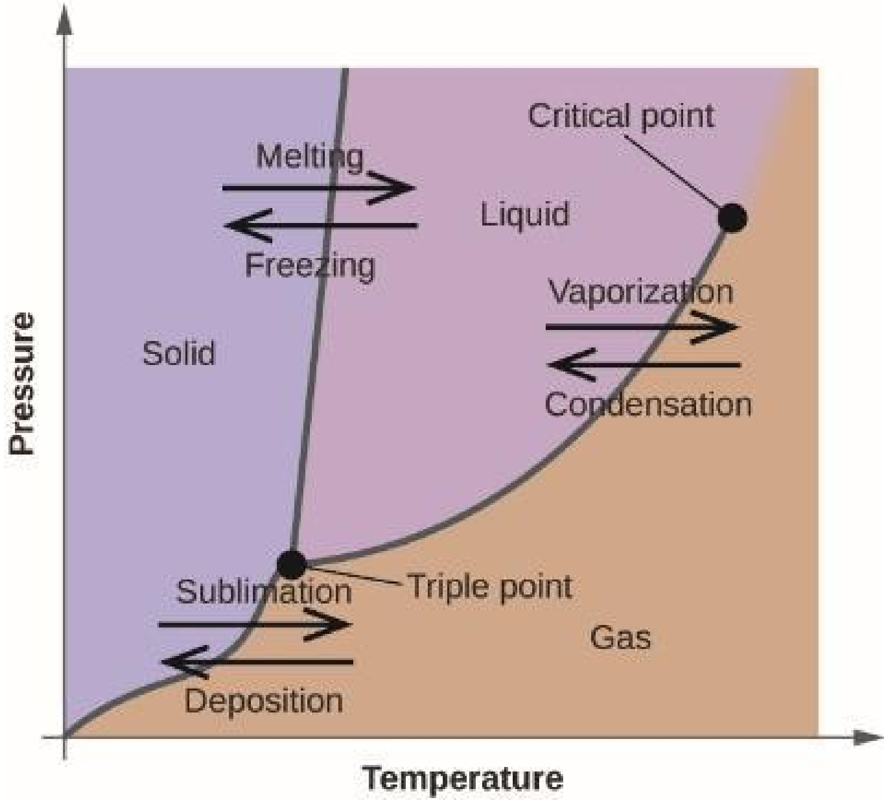
Mott Community College

Phase Transitions and Temperature

- During phase changes, the temperature remains constant.
- If energy is being absorbed in the phase change, then all this energy is being used to break the intermolecular forces, not increase the temperature of the substance.
- If energy is being released during the phase change, then all this released energy is coming from the formation of intermolecular forces, not from the decrease in temperature of the substance.

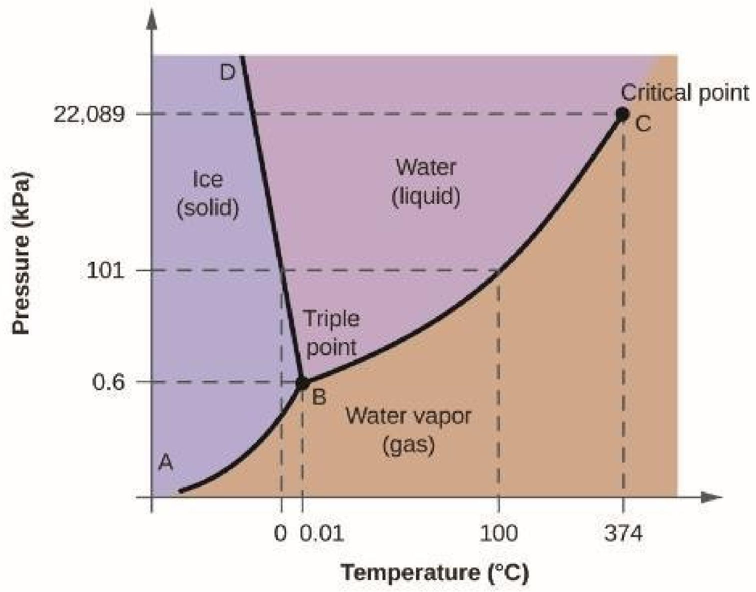
Phase Diagrams

- The physical state of a substance and its phase-transition temperatures are represented graphically in a phase diagram.



Important Points

- **Triple Point** is the temperature and pressure where solid, liquid and gas exist in equilibrium.
- **The Critical Point** is the temperature and pressure above which a substance can exist as a **Supercritical Fluid**.
- **Supercritical Fluids** have properties similar to both a gas and a liquid.



Critical Temperature and Pressure

- Above the **Critical Temperature** a gas cannot be liquefied no matter the the pressure.
- The pressure required to liquefy a gas at its critical temperature is the **Critical Pressure.**

Substance	Critical Temperature (°C)	Critical Pressure (kPa)
hydrogen	–240.0	1300
nitrogen	–147.2	3400
oxygen	–118.9	5000
carbon dioxide	31.1	7400
ammonia	132.4	11,300
sulfur dioxide	157.2	7800
water	374.0	22,000

Supercritical Fluid Properties

- SFs expand to fill their containers.
- Their density is much greater than typical gases.
- They are capable of dissolving nonvolatile solutes.
- They have essentially no surface tension and very low viscosity
- They are extremely useful solvents for many applications.
- Other properties depend on the identity of the fluid. For instance, supercritical water has a low dielectric constant.

Conclusions

- Phase diagrams are useful representations for visualizing the phase of a substance at a specific temperature and pressure.
- The **Triple Point** and **Critical Point** can be found on a phase diagram.
- Supercritical fluids have important properties that differ from solids, liquids, and gases.

Section 10.5
The Solid State of Matter

Michael Stogsdill

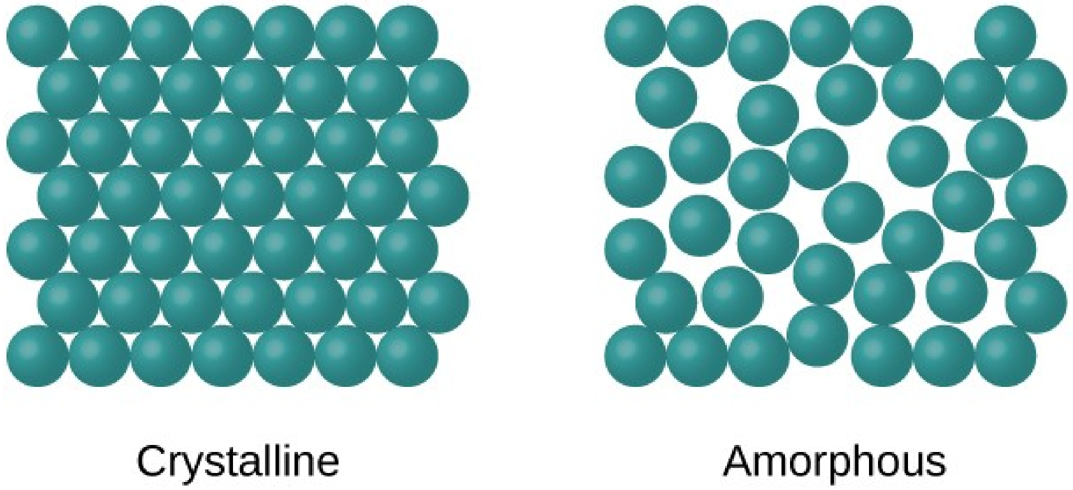
Mott Community College

Learning Objectives

- Define and describe the bonding and properties of ionic, molecular, metallic, and covalent network crystalline solids
- Describe the main types of crystalline solids: ionic solids, metallic solids, covalent network solids, and molecular solids
- Explain the ways in which crystal defects can occur in a solid

Crystalline vs. Amorphous

- **Crystalline Solids** have atoms, ions, or molecules that are arranged in a definite repeating pattern.
- **Amorphous Solids (noncrystalline solids or glasses)** have particles that lack an ordered internal structure and are randomly arranged.

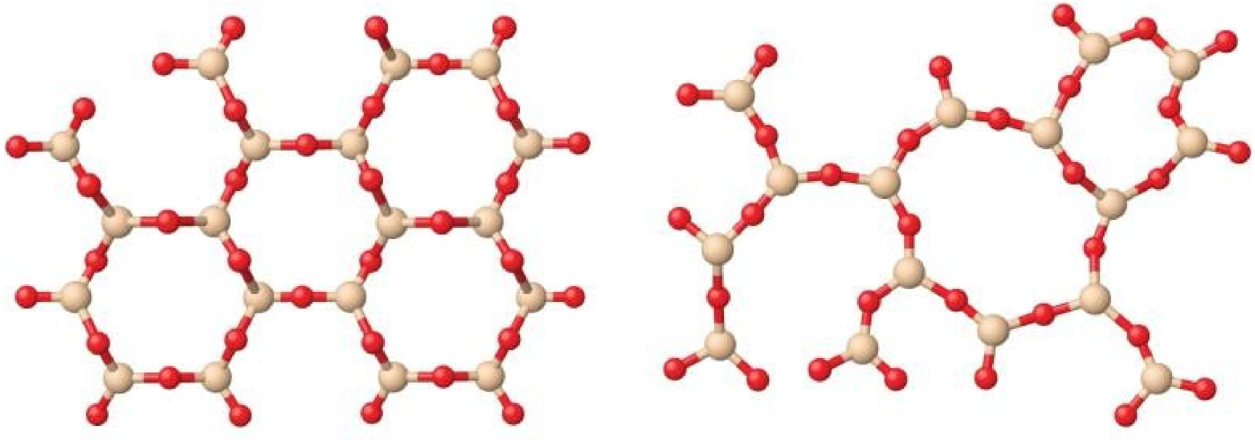


Why One over Another?

- Metals and ionic compounds typically form ordered, crystalline solids.
- Substances that consist of large molecules, or a mixture of molecules whose movements are more restricted, often form amorphous solids.
- Sometimes it depends on the conditions under which the solid formed.

Silicon Dioxide

- (a) Silicon dioxide, SiO_2, is abundant in nature as one of several crystalline forms of the mineral quartz.
- (b) Rapid cooling of molten SiO_2 yields an amorphous solid known as “fused silica”.



(a) (b)

Ionic Solids

- **Ionic Solids** are composed of positive and negative ions that are held together by electrostatic attractions.
- Ionic Interactions are often strong than covalent interactions.
  - Ionic solids will have higher melting points than molecular solids.
- They are hard and brittle.
- Ionic solids do not conduct electricity but molten ionic liquids do.

Ionic Solids

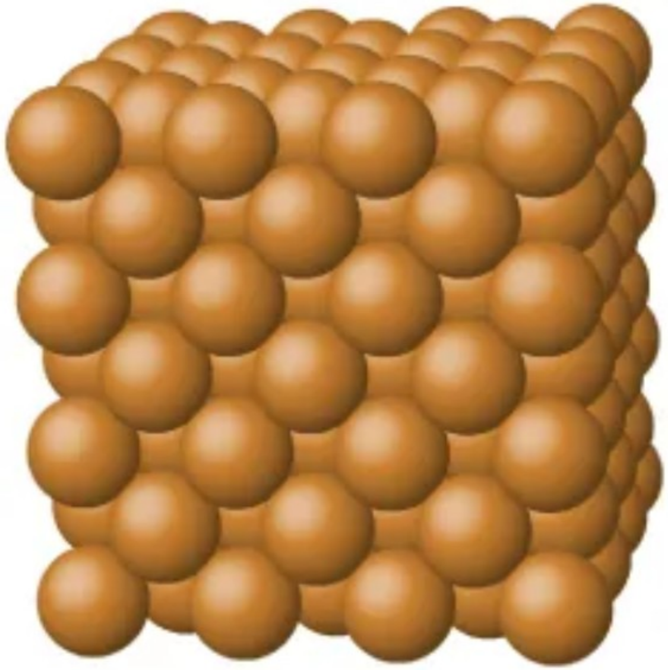
- These solids are usually formed from the the reaction of a metal with a nonmetal.
- Sodium chloride is an ionic solid.

Metallic Solids

- **Metallic Solids** are crystalline structures formed from metal atoms.
- Atoms are held together by a unique force known as **metallic bonding** that gives rise to many useful and varied bulk properties.
- They exhibit high thermal conductivity, electrical conductivity, metallic luster, malleability, and ductility.
- They typically have lower melting points than ionic solids

Metallic Solids

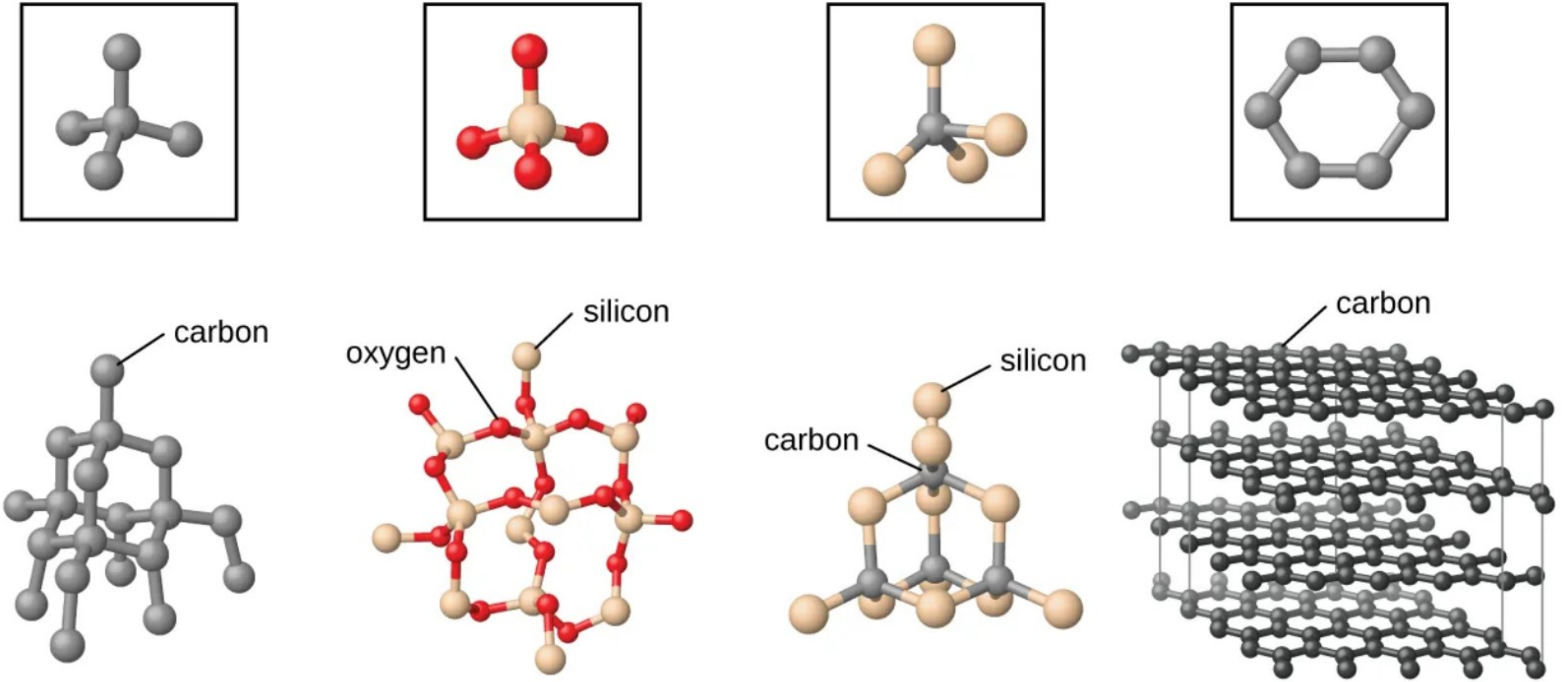
- Metallic crystals are often described as a uniform distribution of atomic nuclei within a “sea” of delocalized electrons.
- Copper is a metallic solid.



Covalent Network Solid

- **Covalent network solids** are held together by interconnecting network of covalent bonds.
- To break or to melt a covalent network solid, covalent bonds must be broken.
- They are typically hard and strong materials with high melting points.
- Diamond is one of the hardest substances known and melts above 3500 °C.

Covalent Network Solids

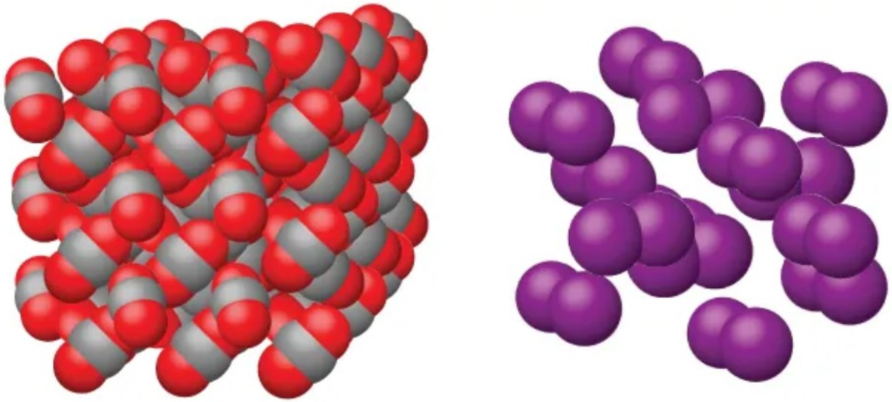


Molecular Solids

- **Molecular solids** are composed of neutral molecules.
- The strengths of the attractive forces between the units present in different crystals vary widely, as indicated by the melting points of the crystals.
- As the IMFs of the molecules increase in the strength the melting points rise.
  - Small molecule dispersive < Large Molecule Dispersive < Dipole-Dipole

Molecular Solids

- Carbon dioxide (CO_2) consists of small, nonpolar molecules and forms a molecular solid with a melting point of −78 °C. Iodine (I_2) consists of larger, nonpolar molecules and forms a molecular solid that melts at 114 °C.



carbon dioxide iodine

Properties of Solids

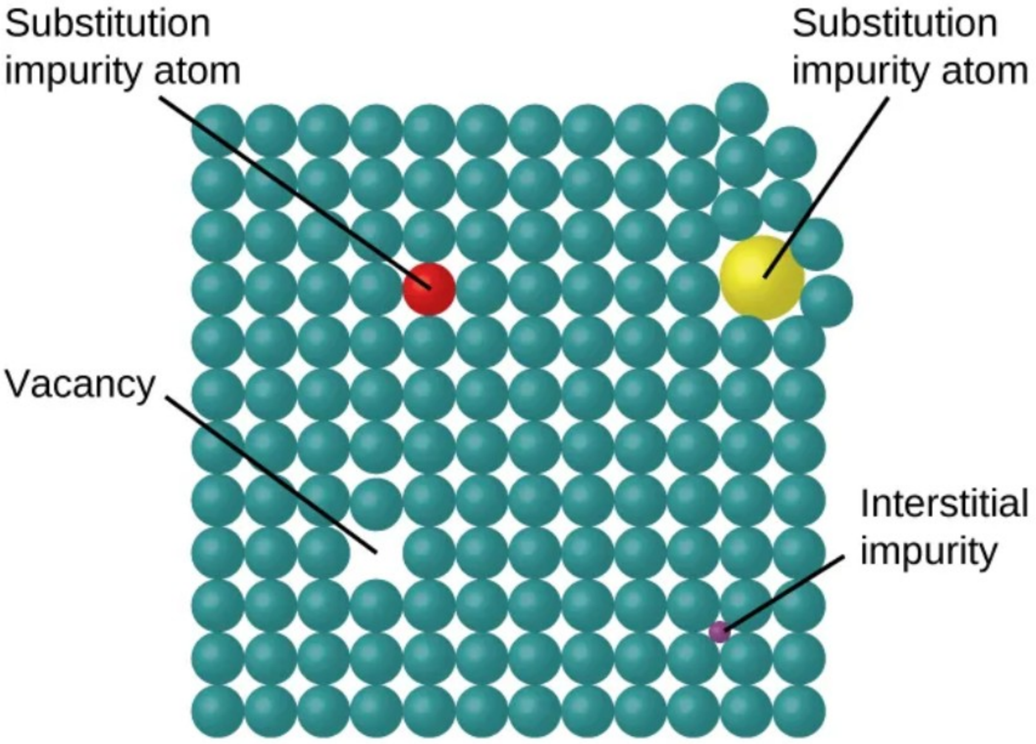
Type of Solid	Type of Particles	Type of Attractions	Properties	Examples
ionic	ions	ionic bonds	hard, brittle, conducts electricity as a liquid but not as a solid, high to very high melting points	$NaCl$, Al_2O_3
metallic	atoms of electropositive elements	metallic bonds	shiny, malleable, ductile, conducts heat and electricity well, variable hardness and melting temperature	Cu, Fe, Ti, Pb, U
covalent network	atoms of electronegative elements	covalent bonds	very hard, not conductive, very high melting points	C (diamond), SiO_2, SiC
molecular	molecules (or atoms)	IMFs	variable hardness, variable brittleness, not conductive, low melting points	H_2O, CO_2, I_2, $C_{12}H_{22}O_{11}$

Crystal Defects

- Occasional defects may occur in the pattern of a crystalline solid.
- **Vacancies** are defects that occur when positions that should contain atoms or ions are vacant.
- Sometimes atoms may occupy “out of place” locations called **interstitial sites**.
- Sometimes impurities are trapped in the crystal lattice.
- The practice of intentionally including impurities to produce desired properties is called **doping**.

Crystal Defects

- Types of crystal defects include vacancies, interstitial atoms, and substitutions impurities.



Section 10.6
Lattice Structures in Crystalline Solids

Michael Stogsdill

Mott Community College

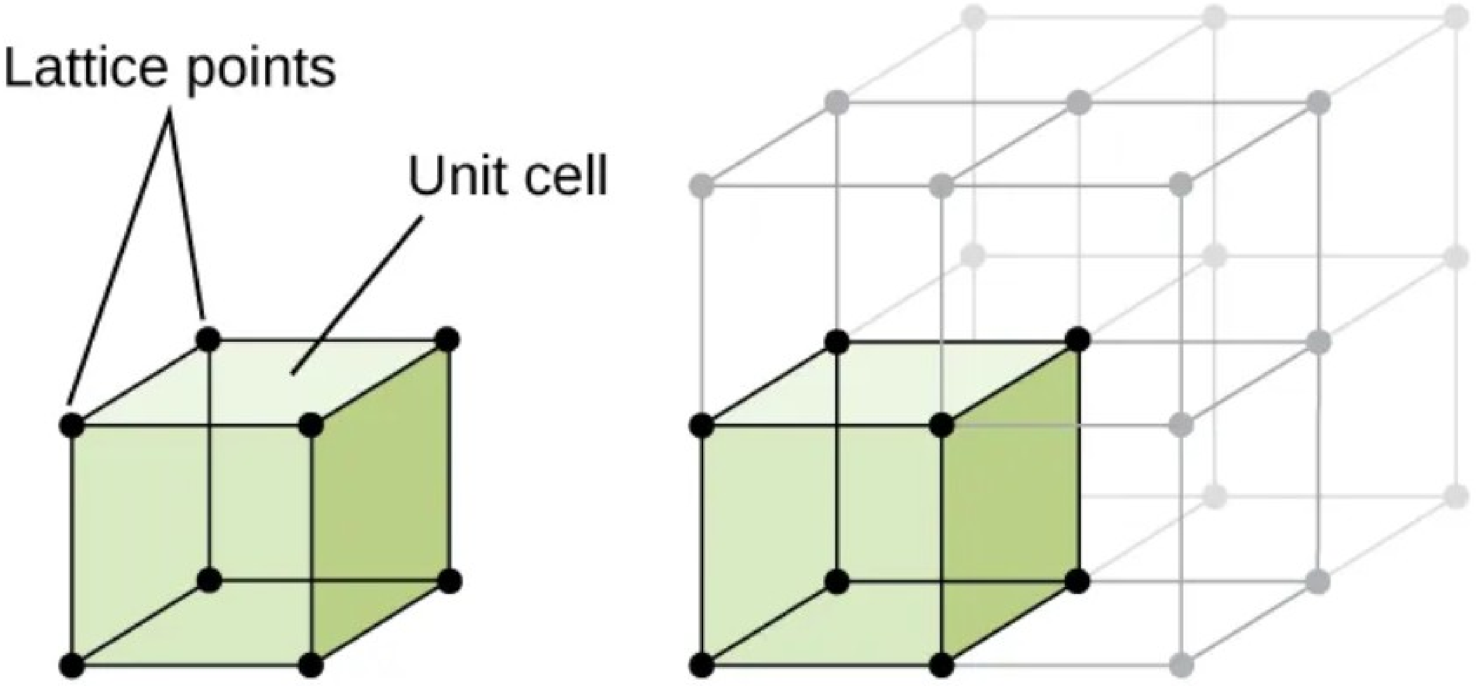
Learning Objectives

- Describe the arrangement of atoms and ions in crystalline structures
- Compute ionic radii using unit cell dimensions
- Explain the use of X-ray diffraction measurements in determining crystalline structures

Unit Cells

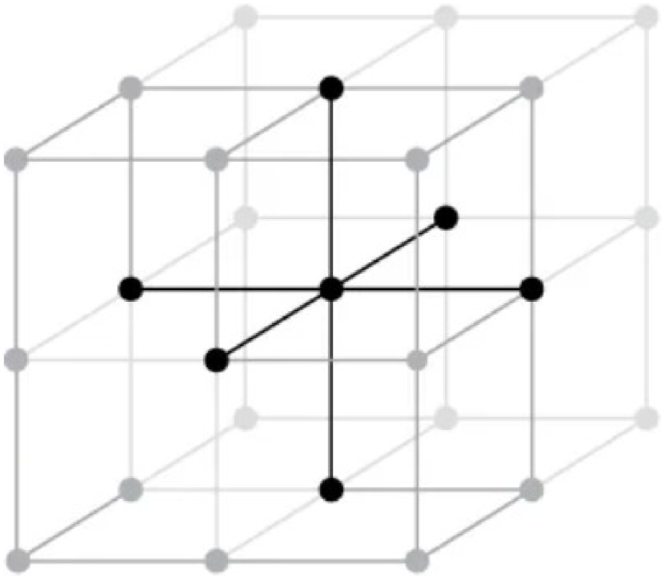
- Crystalline solids are closely packed atoms in a repeating pattern.
- The **unit cell** consists of lattice points that represent the locations of atoms or ions.
- The unit cell is the smallest repeating unit of the overall structure.
- Elements and compounds that crystallize with the same unit cell structure are called **Isomorphous**.

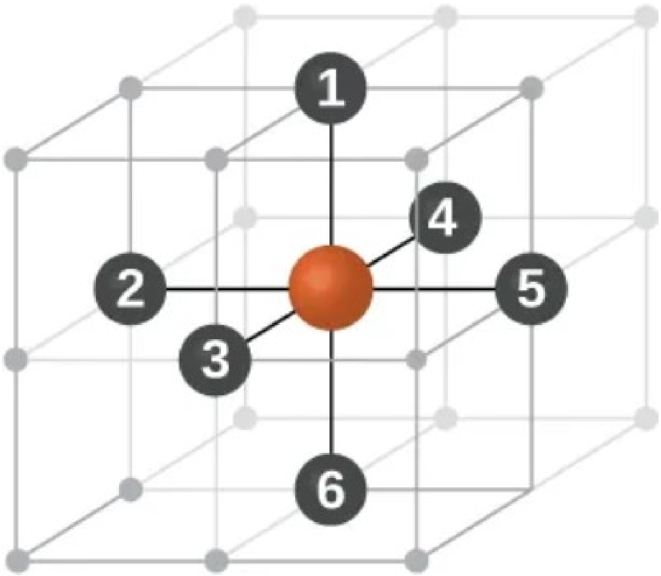
Unit Cells



Coordination Number

- The number of other particles that each particle in a crystalline solid contacts is known as its **Coordination Number**.



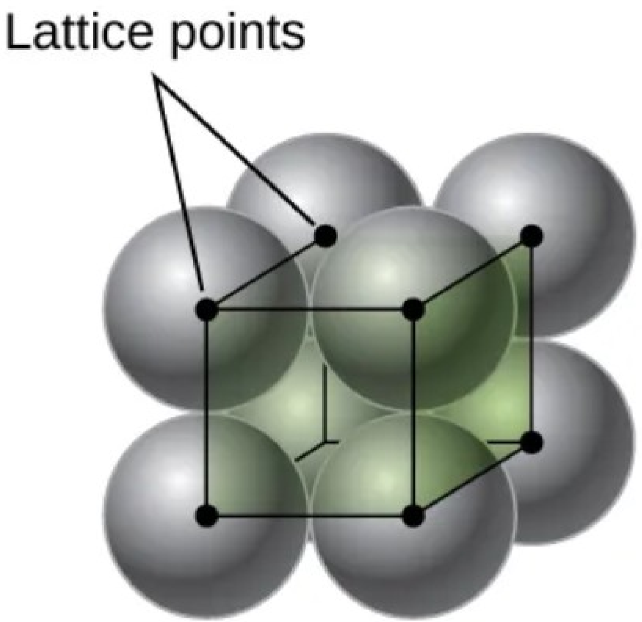


Structures of Metals

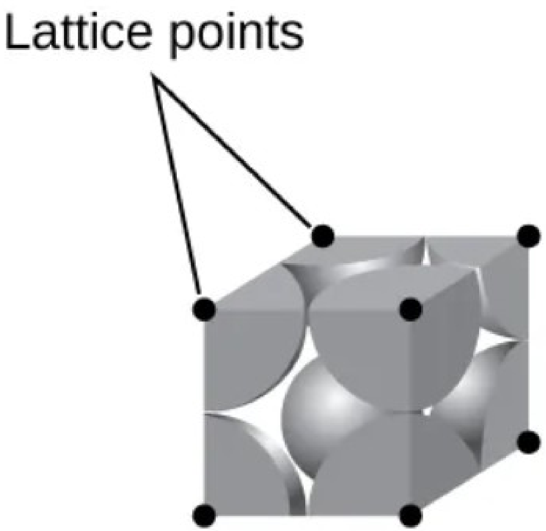
Simple Cubic Structure

- **Simple Cubic Unit Cells** have the atoms lying on the corners of a cube.
- This is an inefficient method of packing atoms. Only about 52% of the overall volume of the unit cell is filled.
- This arrangement has a coordination number of six. Each cell only holds one atom.

Simple Cubic Unit Cell



Simple cubic lattice cell

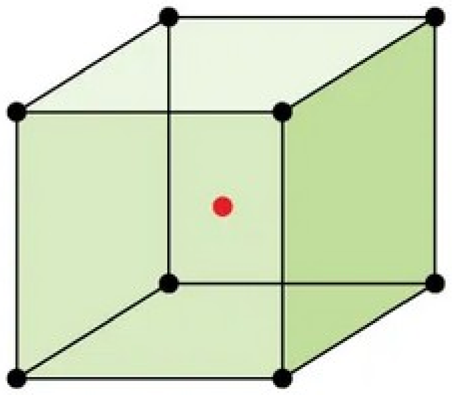


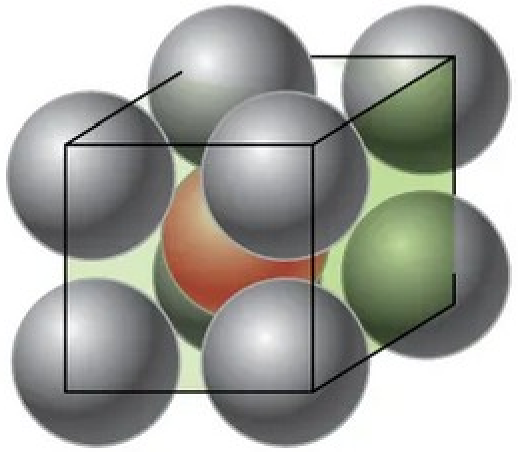
8 corners

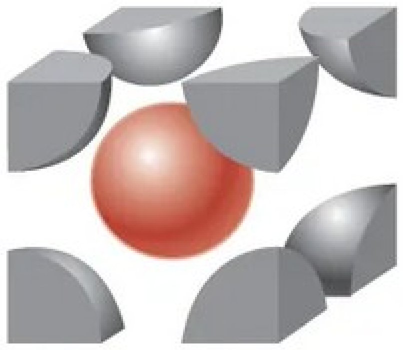
Body-Centered Cubic Structure

- **Body-Centered Cubic (BCC) Solids** have atoms at each corner and an atom in the center.
- BCC solids have a coordination number of eight.
- Each unit cell contains two atoms.
- 68% of the unit cell’s volume is filled.

Body-centered Cubic Unit Cell



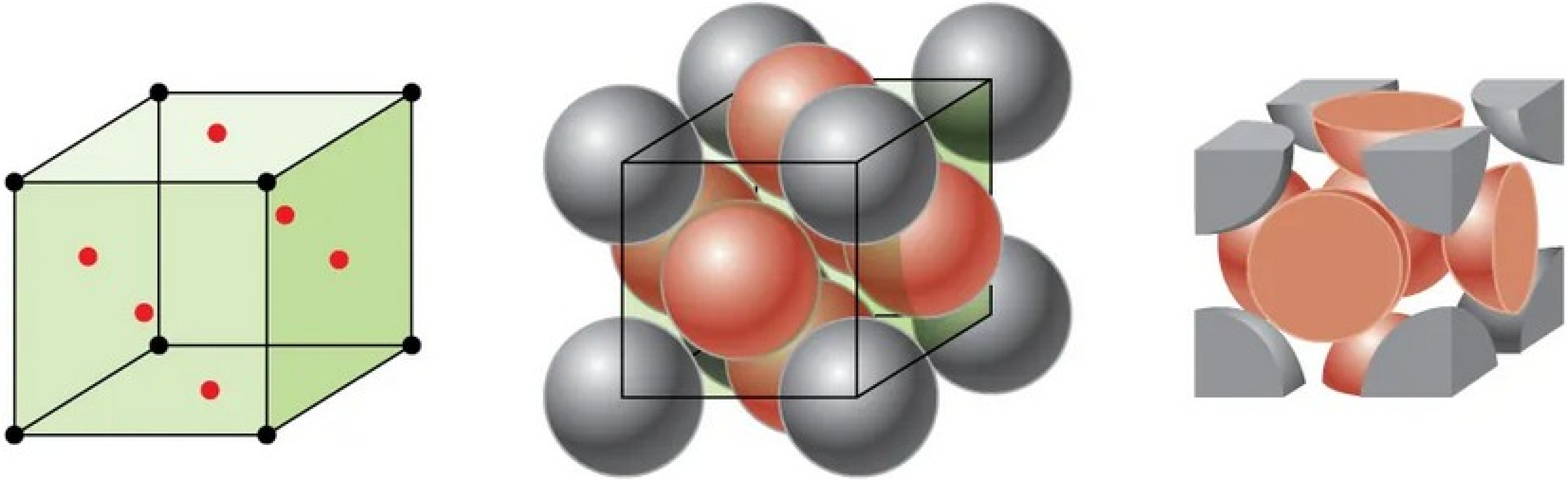




Face-Centered Cubic Structure

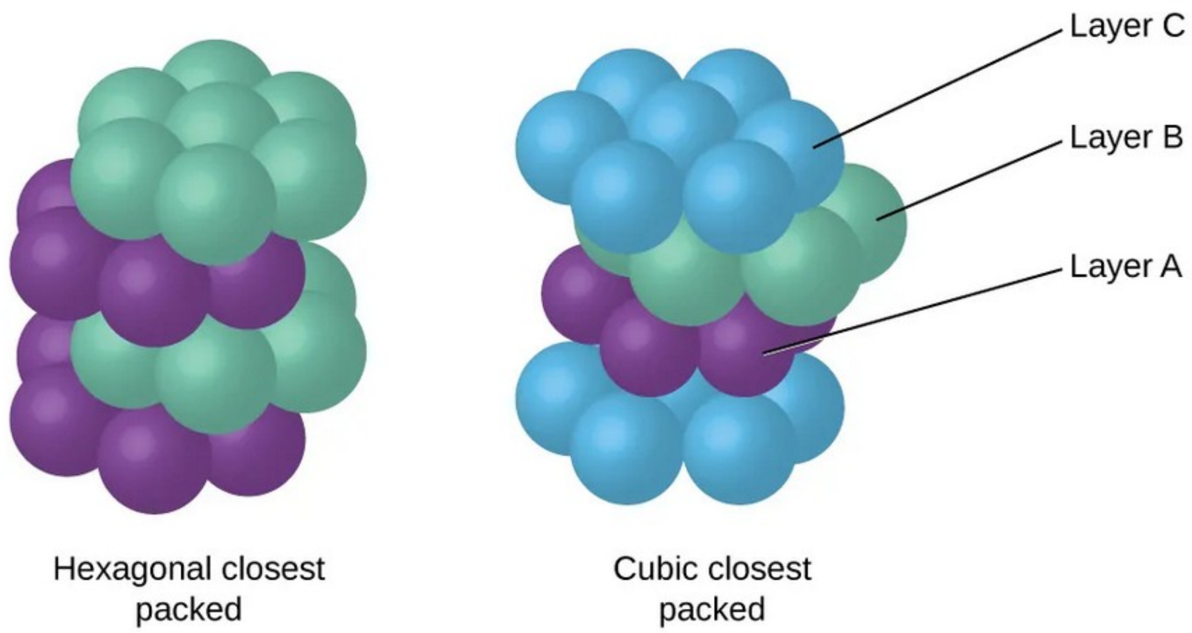
- **Face-Centered Cubic (FCC) Solid** has a cubic unit cell with atoms at all of the corners and at the centers of each face.
- This structure is also called **Cubic Closest Packing (CCP).**
- They have a coordination number of twelve.
- Each Unit cell contains four atoms.
- FCC arrangements pack the cells as efficiently as possible. 74% of the volume is filled.

Face-Centered Unit Cell



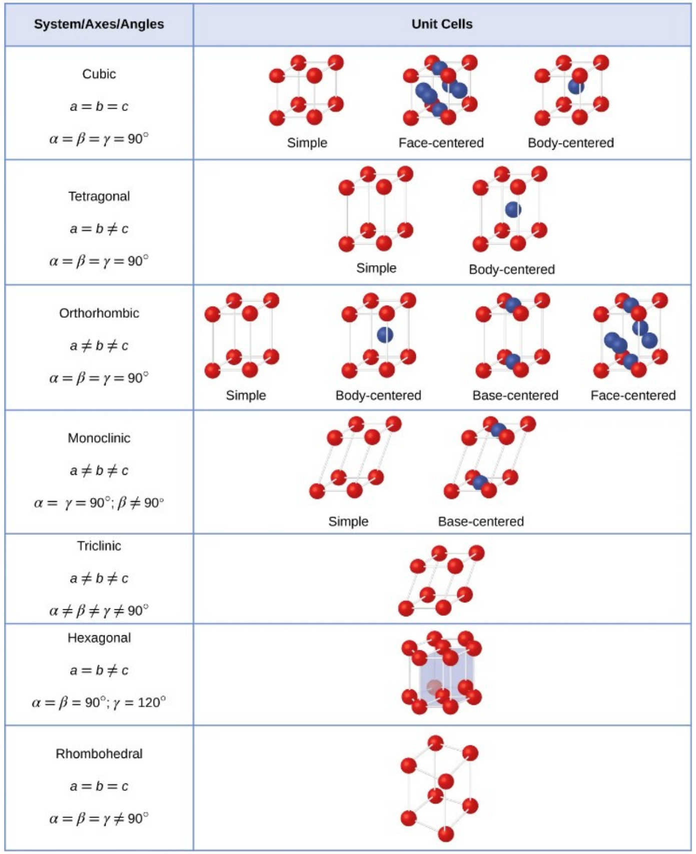
Hexagonal Closest Structure

- **Hexagonal Closest Packing (HCP)** consist of repeating layers of hexagonally arranged atoms.



Other Lattice Systems

- There are seven different lattice systems, some of which have more than one type of lattice, for a total of fourteen different unit cells.



System/Axes/Angles	Unit Cells
Cubic $a = b = c$ $\alpha = \beta = \gamma = 90°$	Simple, Face-centered, Body-centered
Tetragonal $a = b \neq c$ $\alpha = \beta = \gamma = 90°$	Simple, Body-centered
Orthorhombic $a \neq b \neq c$ $\alpha = \beta = \gamma = 90°$	Simple, Body-centered, Base-centered, Face-centered
Monoclinic $a \neq b \neq c$ $\alpha = \gamma = 90°; \beta \neq 90°$	Simple, Base-centered
Triclinic $a \neq b \neq c$ $\alpha \neq \beta \neq \gamma \neq 90°$	
Hexagonal $a = b \neq c$ $\alpha = \beta = 90°; \gamma = 120°$	
Rhombohedral $a = b = c$ $\alpha = \beta = \gamma \neq 90°$	

Structures of Ionic Crystals

Ionic Crystal Structure

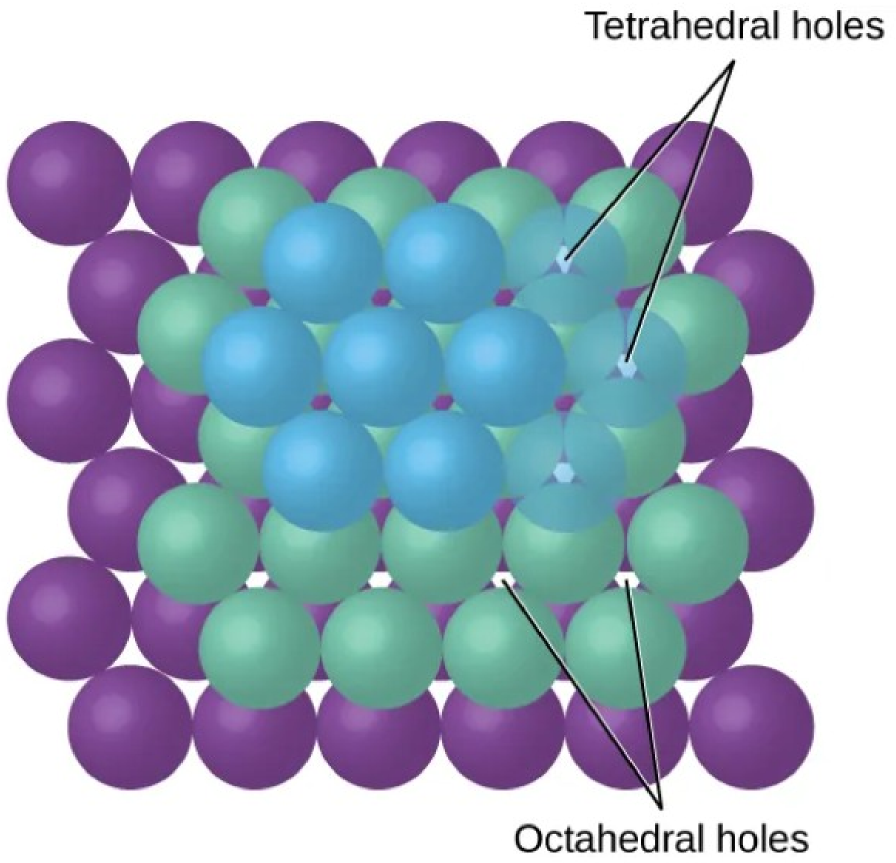
- The packing of ions into a crystal structure is more complex than the packing of metal atoms that are the same size.
- Stable structures for ionic compounds result when
  - Ions of one charge are surrounded by as many ions as possible of the opposite charge
  - Cations and anions are in contact with each other.
- Structures are determined by two principal factors
  - Relative sizes of the ions
  - Ratio of the numbers of positive and negative ions in the compound.

Simple Ionic Structures

- Larger anions are arranged in a closest-packed array.
- Smaller cations commonly occupy one of two types of **holes (or interstices)** between the anions.
- **Tetrahedral Holes** are found at the center of four atoms and are smaller.
- **Octahedral Holes** are found at the center of six atoms and are larger.

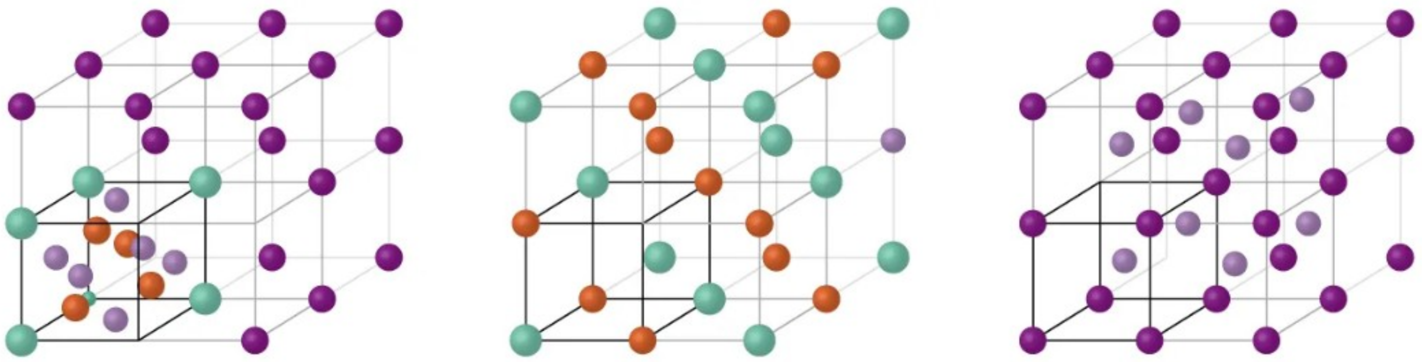
Ionic Holes

- Relatively small cations occupy tetrahedral holes, and larger cations occupy octahedral holes.



Cubic Holes

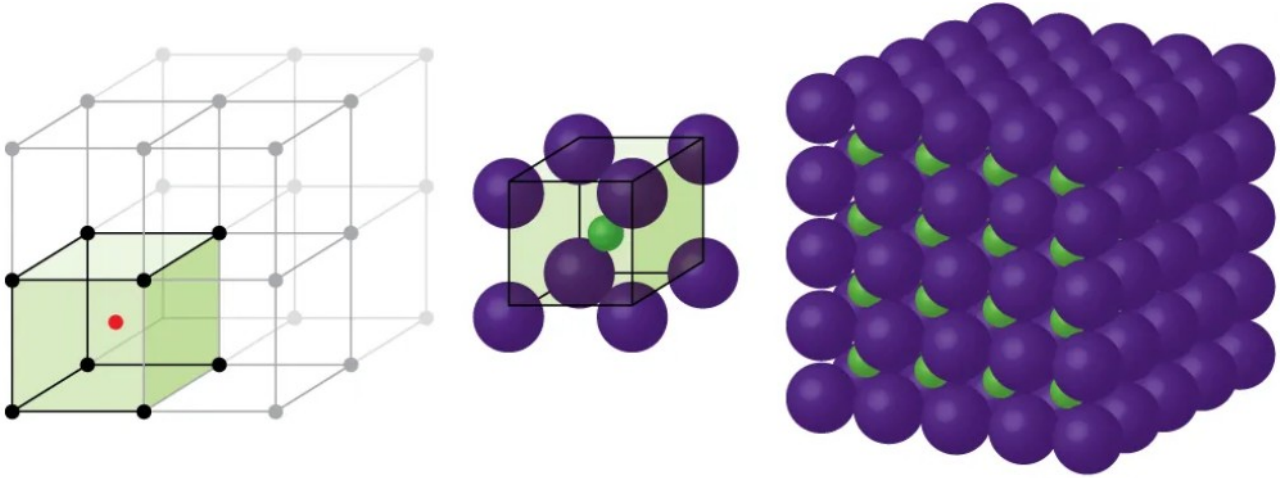
- If the cations are too large to fit into the octahedral holes, the anions may adopt a more open structure, such as a simple cubic array. The larger cations can then occupy the larger cubic holes made possible by the more open spacing.



Tetrahedral hole	Octahedral hole	Cubic hole
Cation radius is about 22.5 to 41.4% of the anion radius	Cation radius is about 41.4 to 73.2% of the anion radius	Cation radius is about 73.2 to 100% of the anion radius

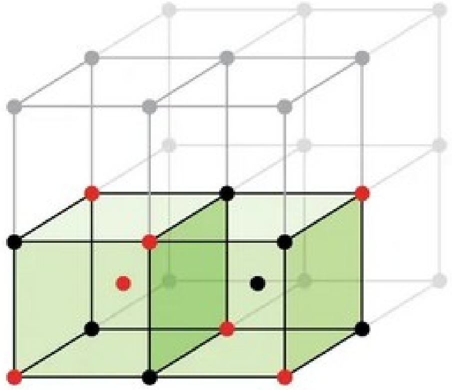
Unit Cells of Ionic Compounds

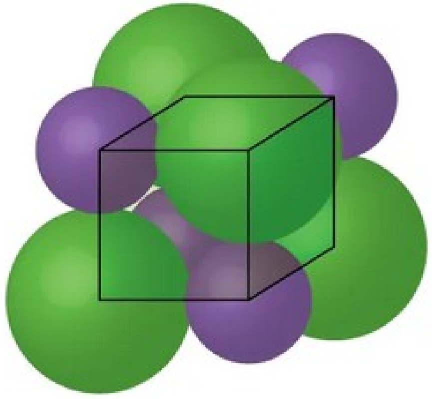
- When an ionic compound is composed of cations and anions of similar size in a 1:1 ratio, it typically forms a simple cubic structure.

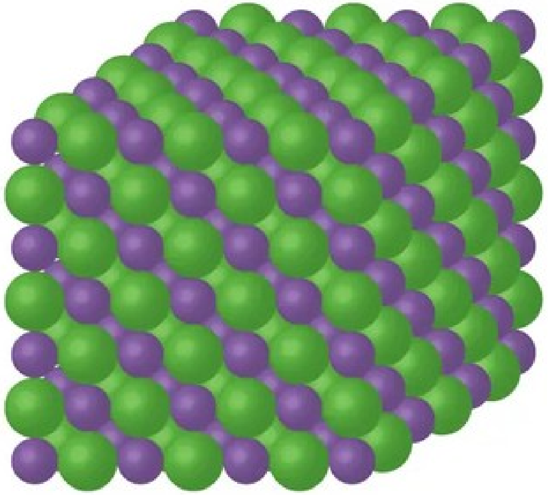


Unit Cells of Ionic Compounds

- When an ionic compound is composed of a 1:1 ratio of cations and anions that differ significantly in size, it typically crystallizes with an FCC unit cell.

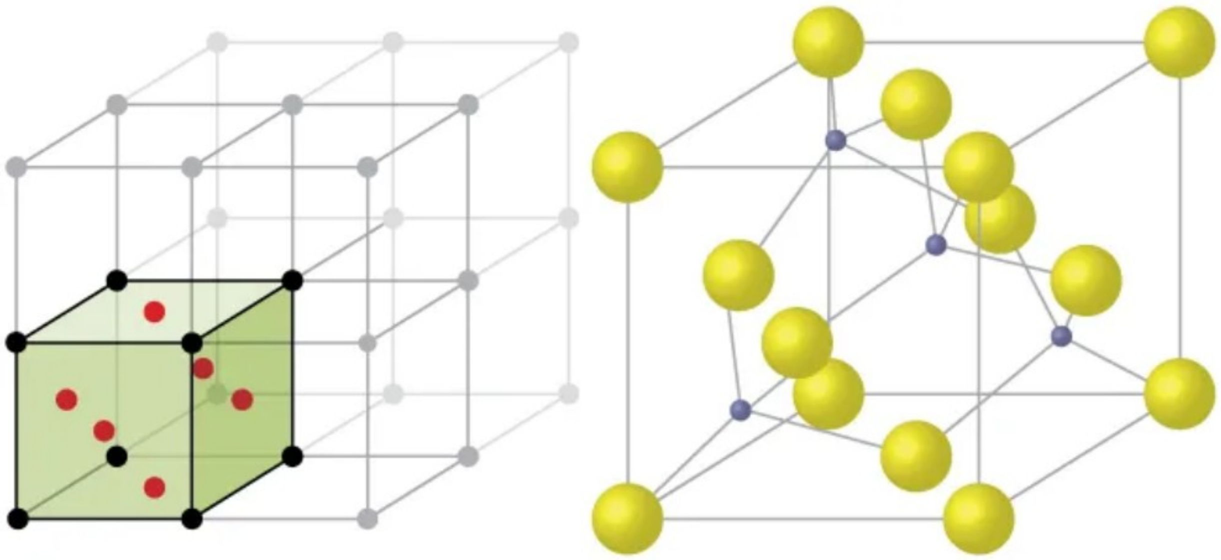






Tetrahedral Holes

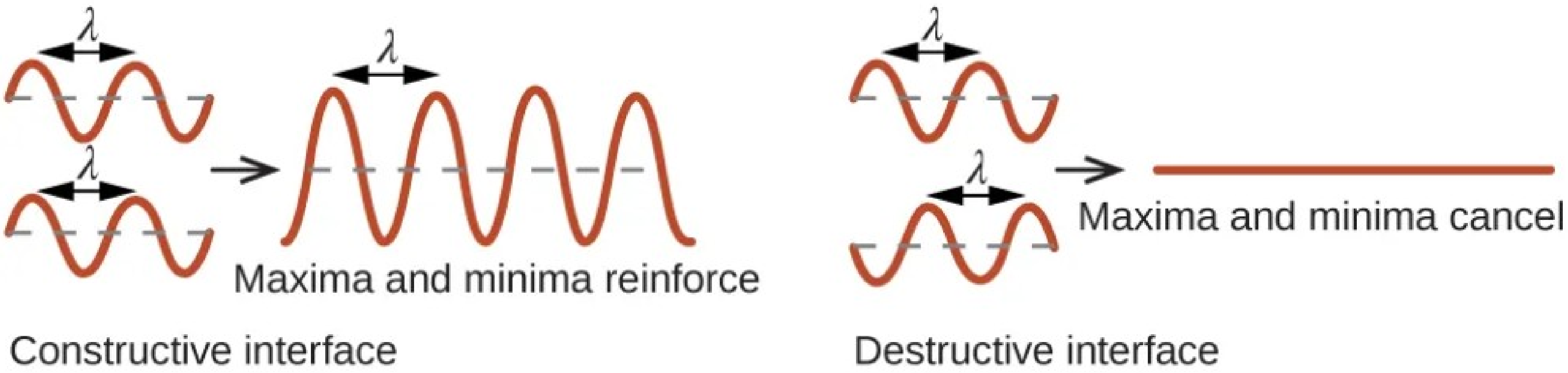
- ZnS forms an FCC unit cell with sulfide ions at the lattice points and much smaller zinc ions occupying half of the tetrahedral holes in the structure.



X-Ray Crystallography

- **X-Ray Crystallography** is a method of determining the crystal structure of a solid by analyzing the diffraction pattern of xrays.
- **Diffraction** is the change in the direction of travel experienced by an electromagnetic wave when it encounters a physical barrier whose dimensions are comparable to those of the wavelength of the light.
- **Interference** is a process by which the waves combine to yield either an increase or a decrease in amplitude depending upon the phase of the two waves.

Constructive vs. Destructive

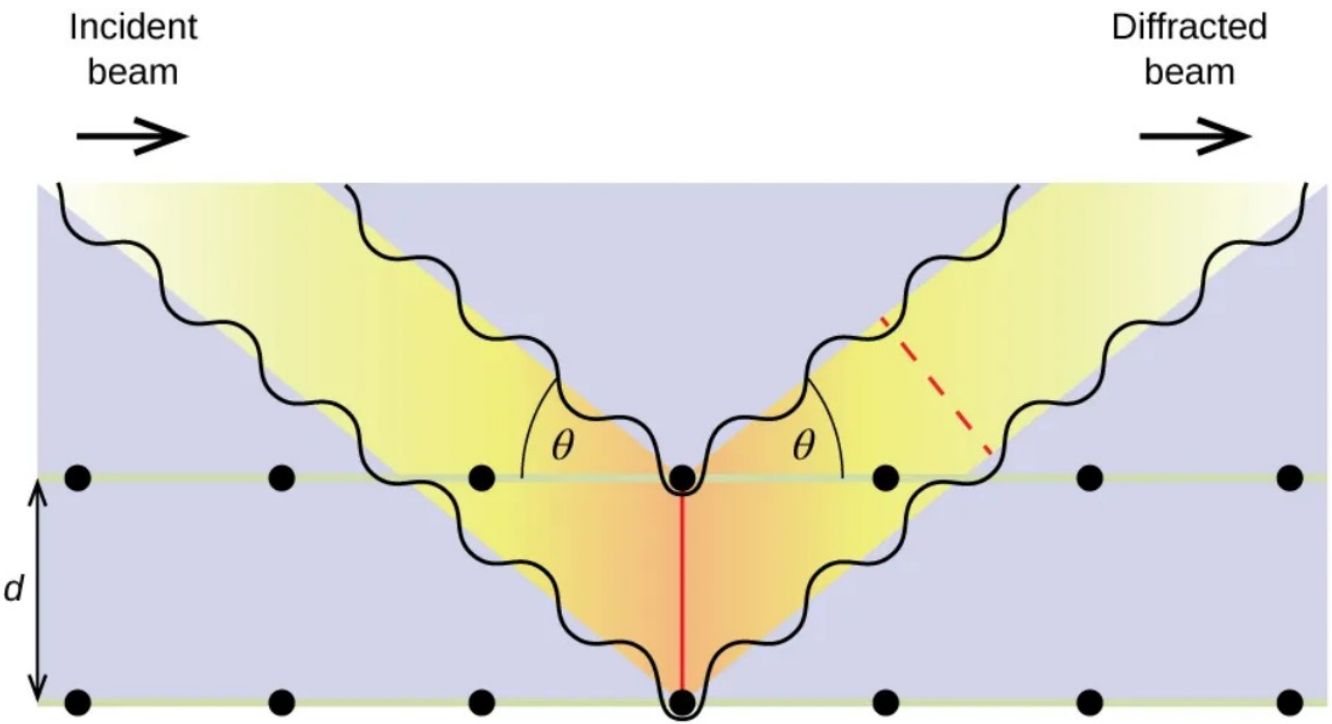


Bragg Equation

- When X-rays of a certain wavelength, λ, are scattered by atoms in adjacent crystal planes separated by a distance, d, they may undergo constructive interference when the difference between the distances traveled by the two waves prior to their combination is an integer factor, n, of the wavelength.

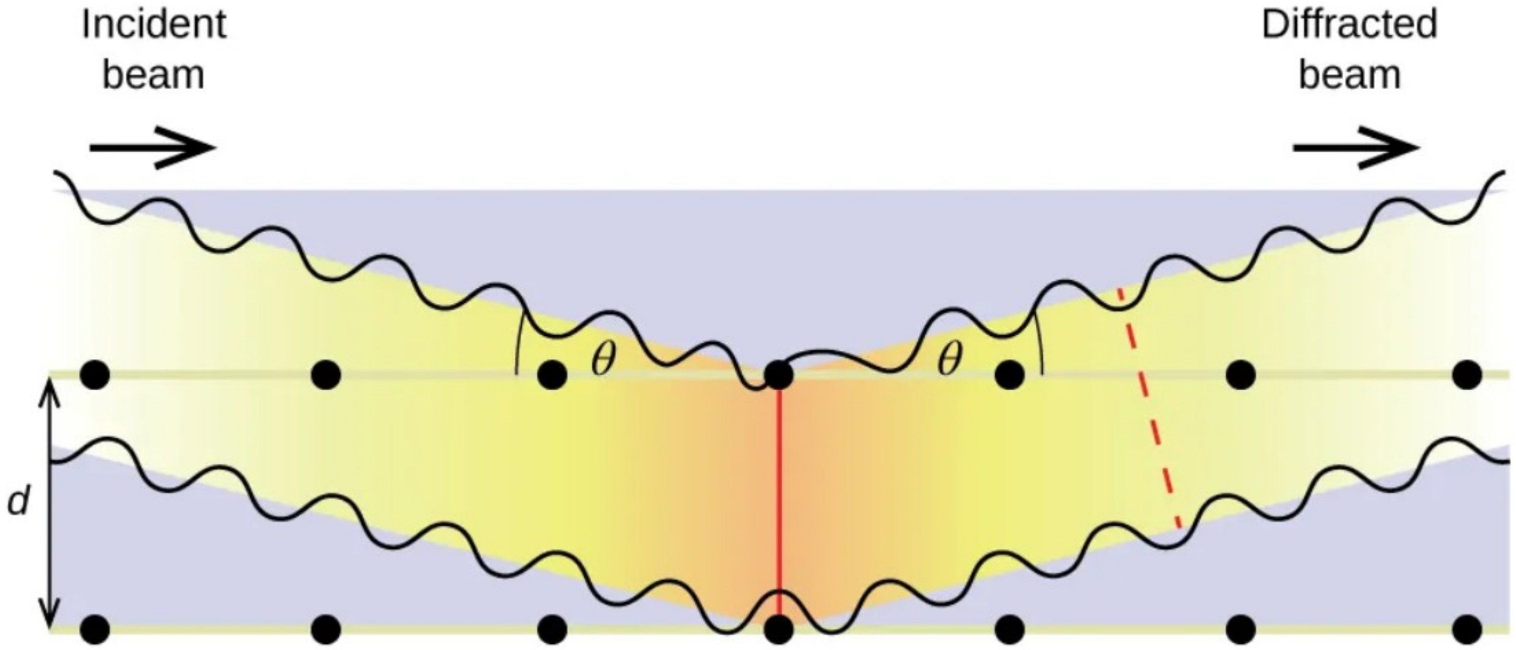
$$n\lambda = 2d\sin\theta$$

Constructive Interference



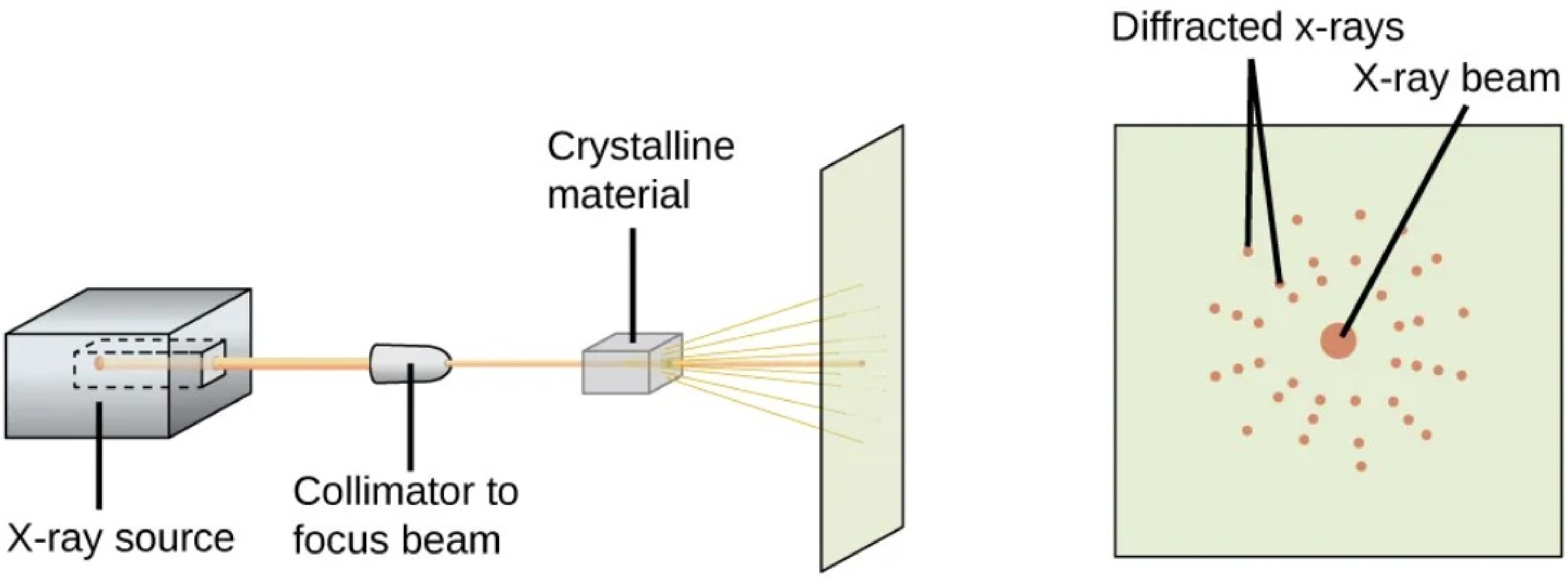
Constructive interference

Destructive Interference



Destructive interference

X-Ray Diffractometer



X-ray diffraction

Imaging surface

X-ray diffraction pattern