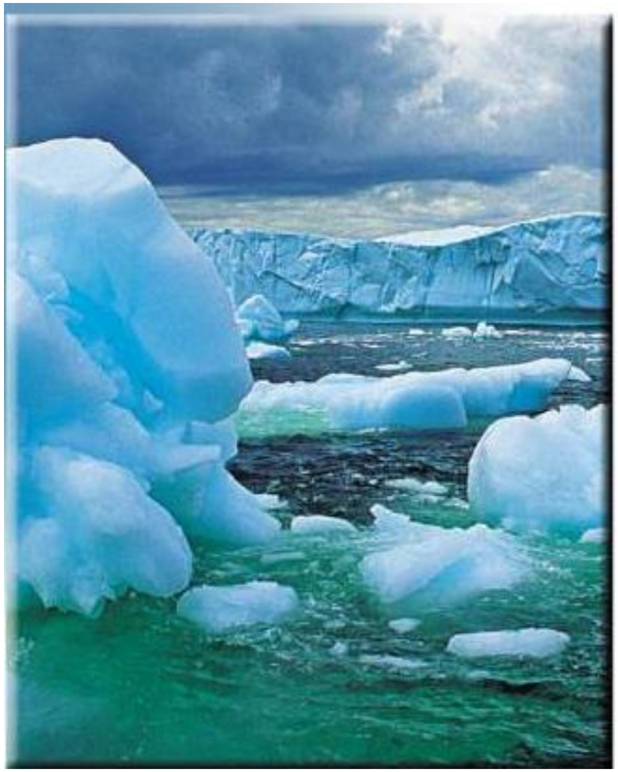


Intermolecular Forces and Liquids and Solids

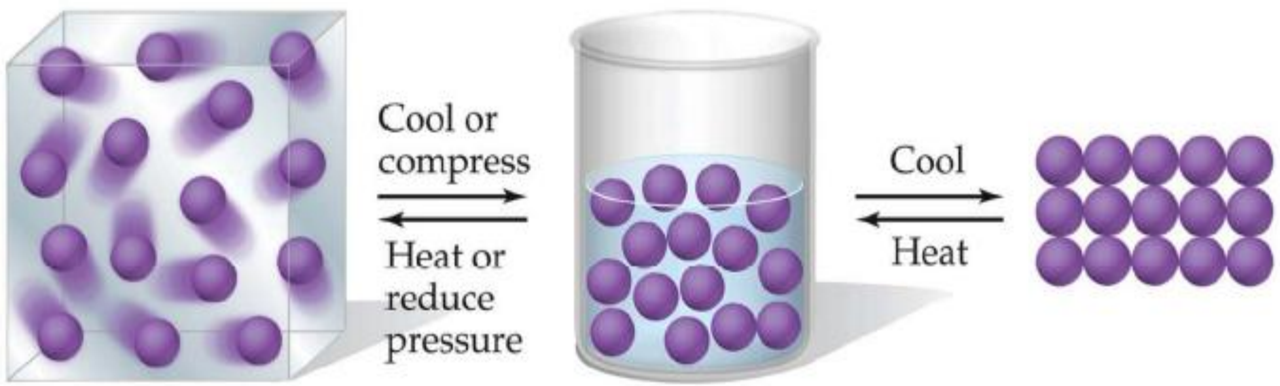
Chapter 11

Based on ppt from Mr. Rapp's AP chemistry website and modified for our needs

http://www.chemistrygeek.com/chem2.htm

States of Matter

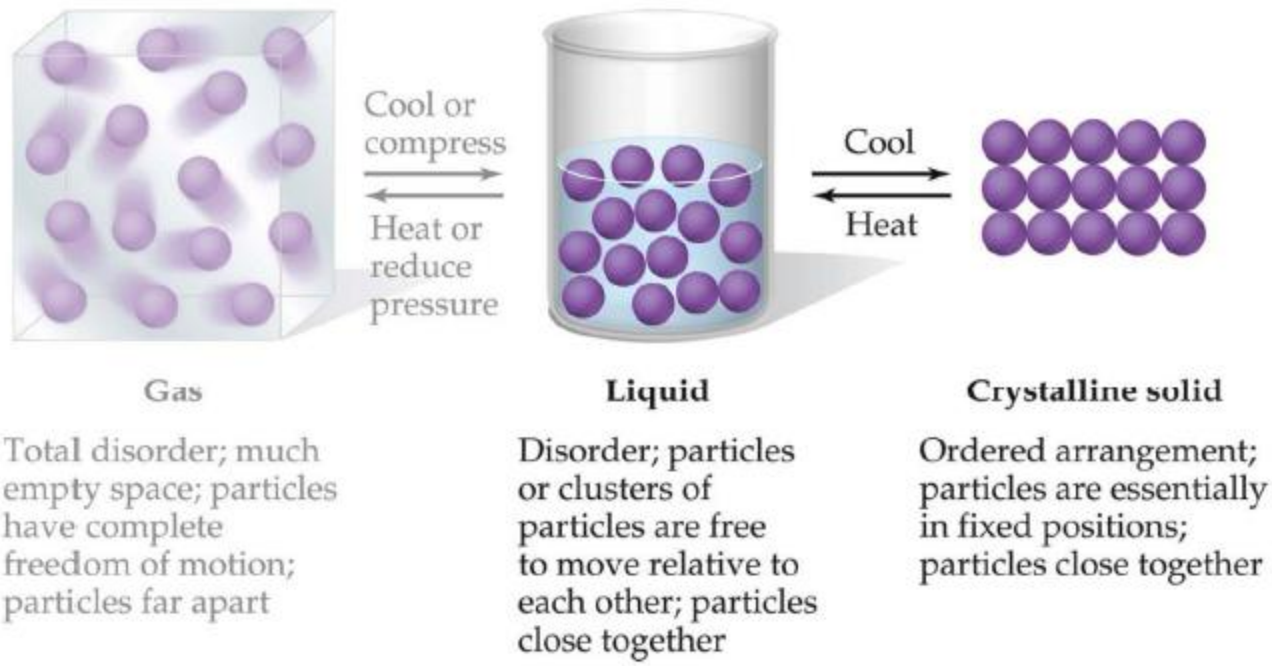
The fundamental difference between states of matter is the distance between particles.



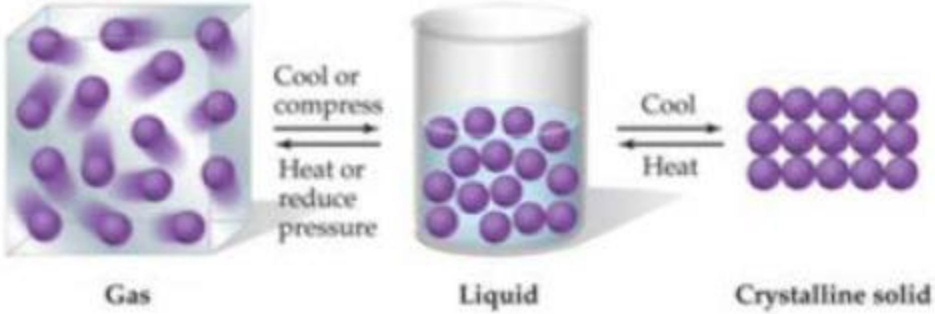
Gas	Liquid	Crystalline solid
Total disorder; much empty space; particles have complete freedom of motion; particles far apart	Disorder; particles or clusters of particles are free to move relative to each other; particles close together	Ordered arrangement; particles are essentially in fixed positions; particles close together

The States of Matter

Because in the solid and liquid states particles are closer together, we refer to them as condensed phases.



The States of Matter



- The state a substance is in at a particular temperature and pressure depends on two antagonistic entities
  - The kinetic energy of the particles
  - The strength of the attractions between the particles

Gas	Assumes both the volume and shape of its container Is compressible Flows readily Diffusion within a gas occurs rapidly
Liquid	Assumes the shape of the portion of the container it occupies Does not expand to fill container Is virtually incompressible Flows readily Diffusion within a liquid occurs slowly
Solid	Retains its own shape and volume Is virtually incompressible Does not flow Diffusion within a solid occurs extremely slowly

A ***phase*** is a homogeneous part of the system in contact with other parts of the system but separated from them by a well-defined boundary.

Intermolecular Forces, Liquids and Solids (Chapter 11)

Resources and Activities

- Textbook - chapter 11 & ppt file
- Online practice quiz from Pearson - chapter 11
- Lab activities
  - Molar mass of a volatile liquid
- POGIL activities:
  - Phase changes
  - Intermolecular Interactions
  - Vapor Pressure Curves
- Comprehensive tutorial and animations on Intermolecular forces:

  http://www.chem.purdue.edu/gchelp/liquids/imf2.html

Chemtour videos from W.W. Norton chapter 10 : intermolecular forces; phase diagrams; (capillary action)

http://www.wwnorton.com/college/chemistry/gilbert2/contents/ch10/studyplan.asp

Chapter 11 Animations from glencoe website for Chang's book:

http://glencoe.mcgraw-hill.com/sites/0023654666/student_view0/chapter11/animations_center.html

Activities and Problem set for chapter 11
(due date________)

TextBook ch. 11 - required for regents (in part), SAT II and AP exams

Lab activities: (determining molar mass of volatile liquids)

POGILS (3)

- Phase changes
- Intermolecular Interactions
- Vapor Pressure Curves

Online practice quiz ch 11 due by______

- Do chapter 11 **GIST** (p. 445, 446, 447, 448, 452, 454, 456, 461, 463) and **Visualizing concepts** problems 11.1-11.6 (6 total) - write out questions and answers & show work. (NB Photocopy of questions is also acceptable)
- Do End of chapter exercises:11.9, .11, .13, .15, .17, .19, .21, .23, .25, .27, .29, .33, .35, .37, .39, .41, .43, .45, .47, .49, .51, .53, .55, .71, .73, .75, .85, .100

Animation on intermolecular forces, liquids, and solids to view in class and at home:

http://www.wwnorton.com/college/chemistry/gilbert2/contents/ch10/studyplan.asp

(Intermolecular forces)

Animations on vapor pressure and unit cells to view in class and at home:

In class preview and then Independent work - students to view animations & interactive activities (4 in total – 1 from Norton and 3 from the Glencoe site for Chang's book) and write summary notes on each. These summaries are to be included in your portfolio

http://glencoe.mcgraw-hill.com/sites/0023654666/student_view0/chapter11/animations_center.html

Intermolecular Forces

Intermolecular forces are attractive forces **between** molecules.

Intramolecular forces hold atoms together in a molecule.

Intermolecular vs Intramolecular

- 41 kJ to vaporize 1 mole of water (**inter**)
- 930 kJ to break all O-H bonds in 1 mole of water (**intra**)

Generally, **inter**molecular forces are much weaker than **intra**molecular forces.

“Measure” of intermolecular force

boiling point

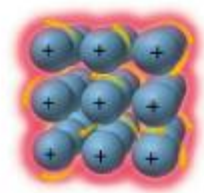
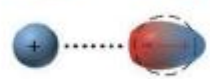
melting point

ΔH_{vap}

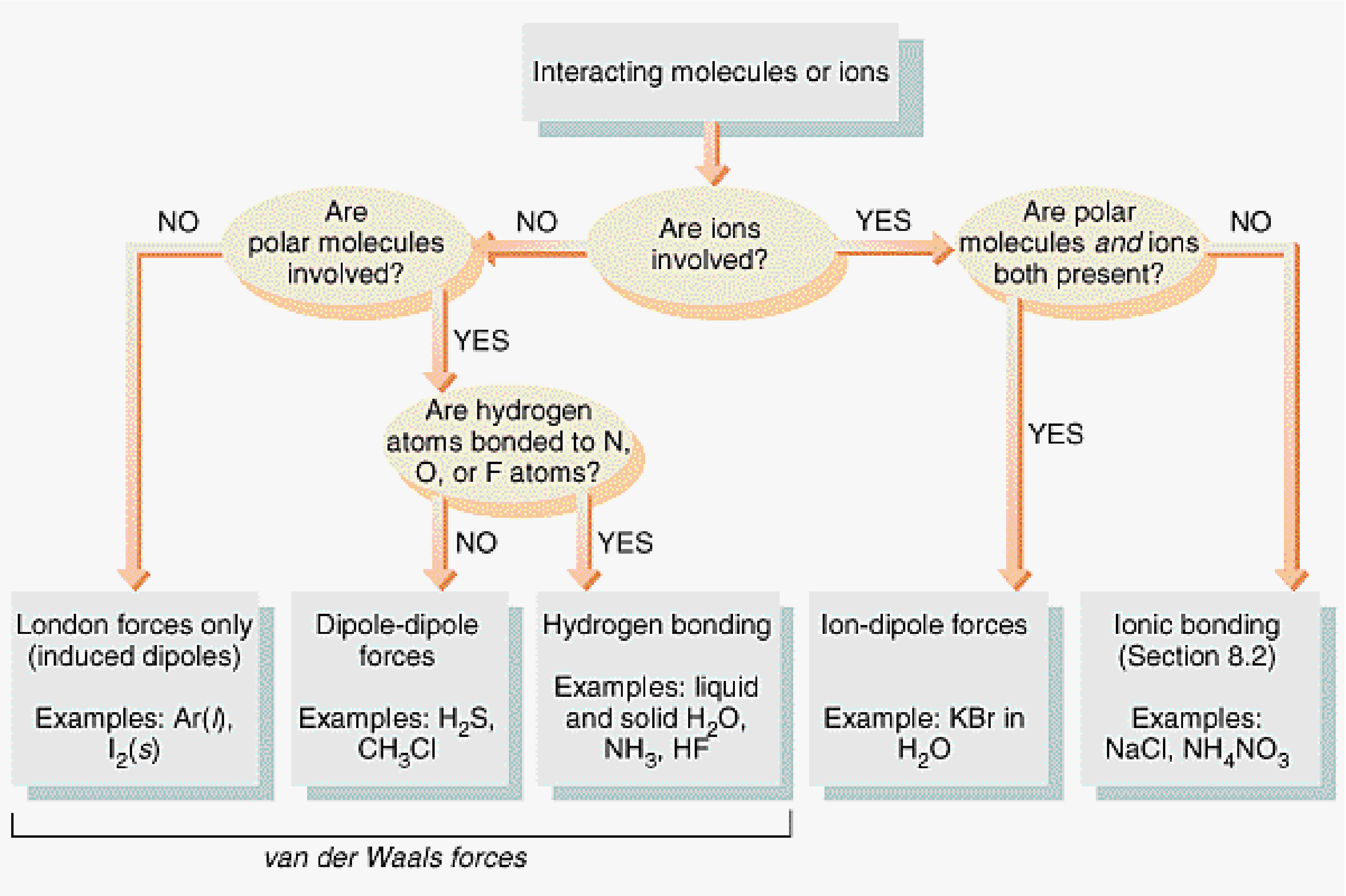
ΔH_{fus}

ΔH_{sub}

Types of Bonding in Crystalline Solids

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
Nonbonding (Intermolecular)				
Ion-dipole		Ion charge–dipole charge	40–600	$Na^+\cdots\cdots O(H)_2$
H bond	δ^- δ^+ δ^- —A—H·······:B—	Polar bond to H–dipole charge (high EN of N, O, F)	10–40	:Ö(H)—H·····:Ö(H)—H
Dipole-dipole		Dipole charges	5–25	I—Cl·····I—Cl
Ion–induced dipole		Ion charge–polarizable e^- cloud	3–15	$Fe^{2+}\cdots\cdots O_2$
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

Molecular Interaction Flowchart



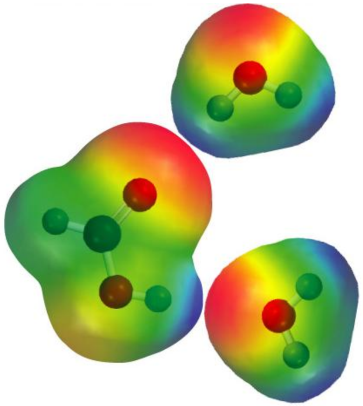
Types of Intermolecular Forces

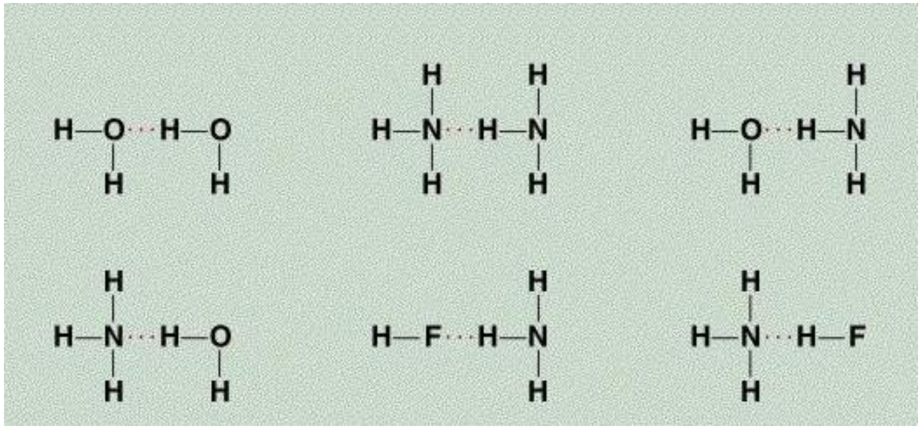
1. Hydrogen Bond (strongest)

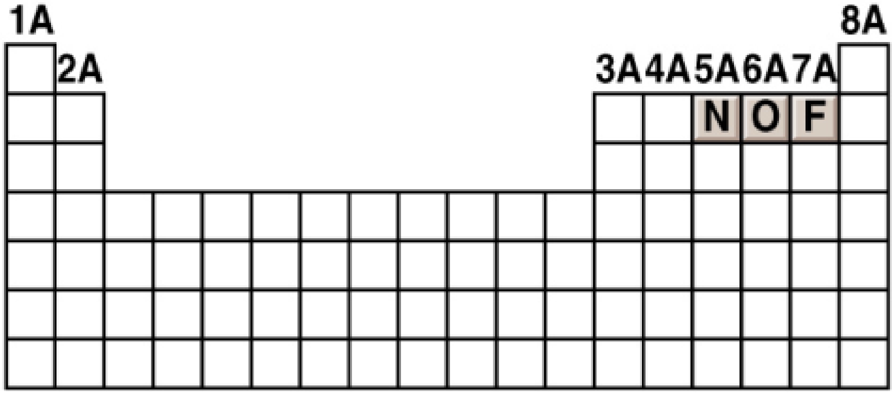
The ***hydrogen bond*** is a special dipole-dipole interaction between the hydrogen atom in a polar N-H, O-H, or F-H bond and an electronegative O, N, or F atom. IT IS NOT A BOND.

A — H··· B or A — H··· A

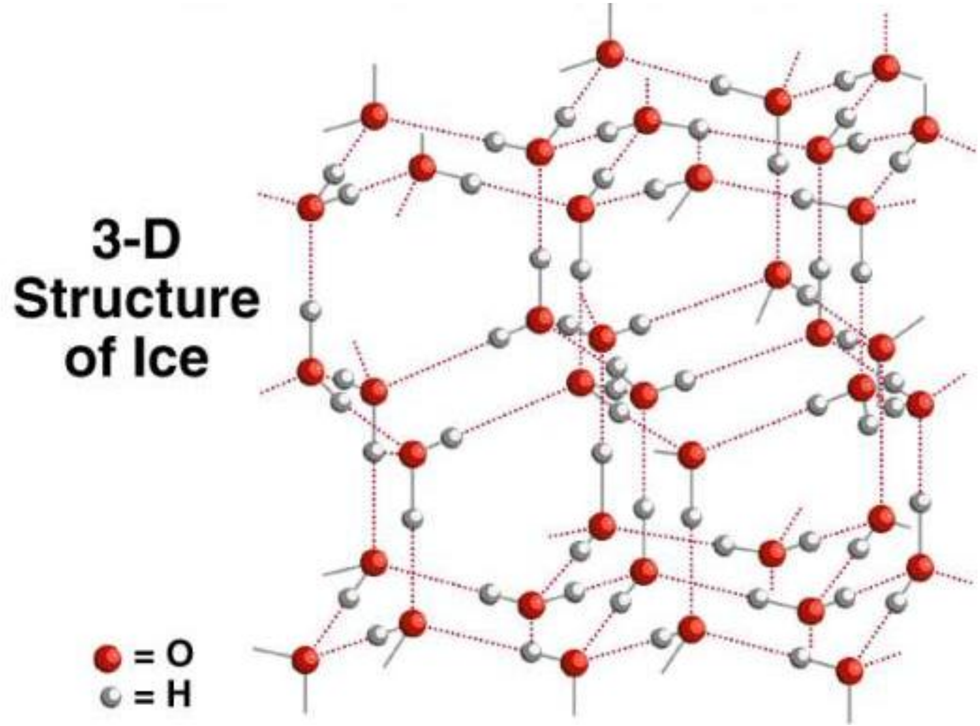
A & B are N, O, or F







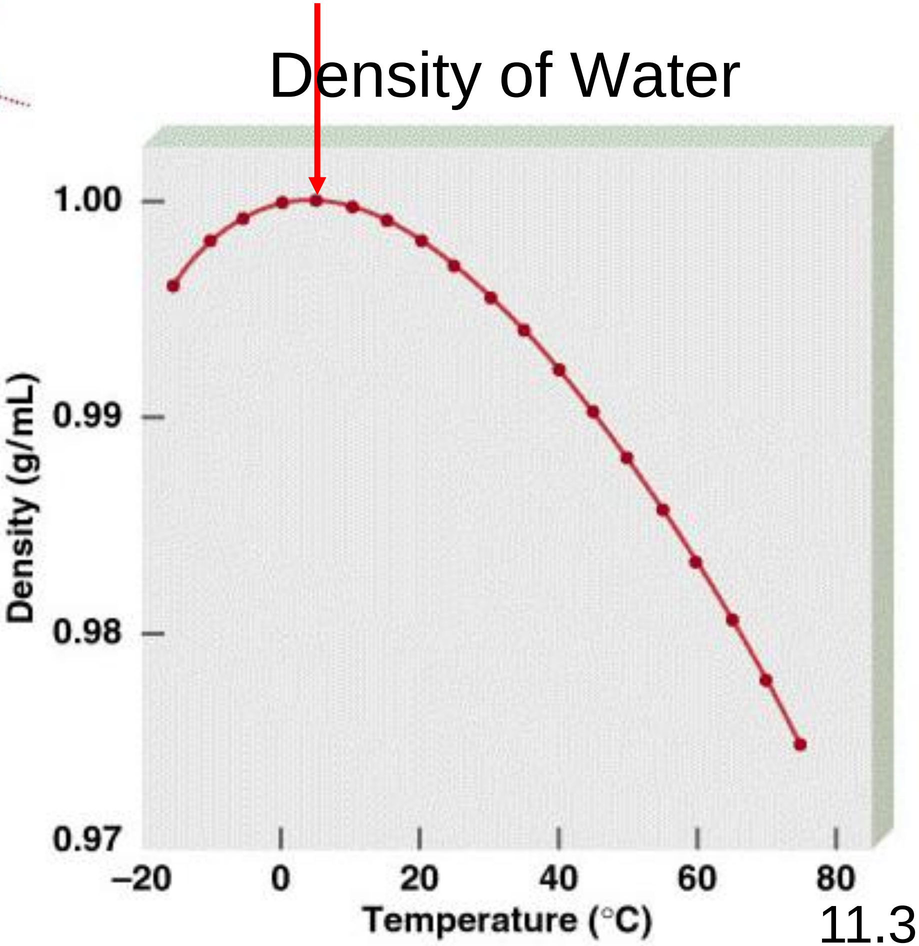
Water is a Unique Substance

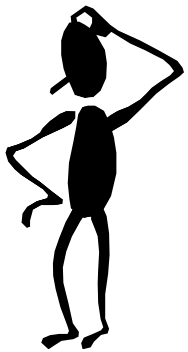


Ice is less dense than water

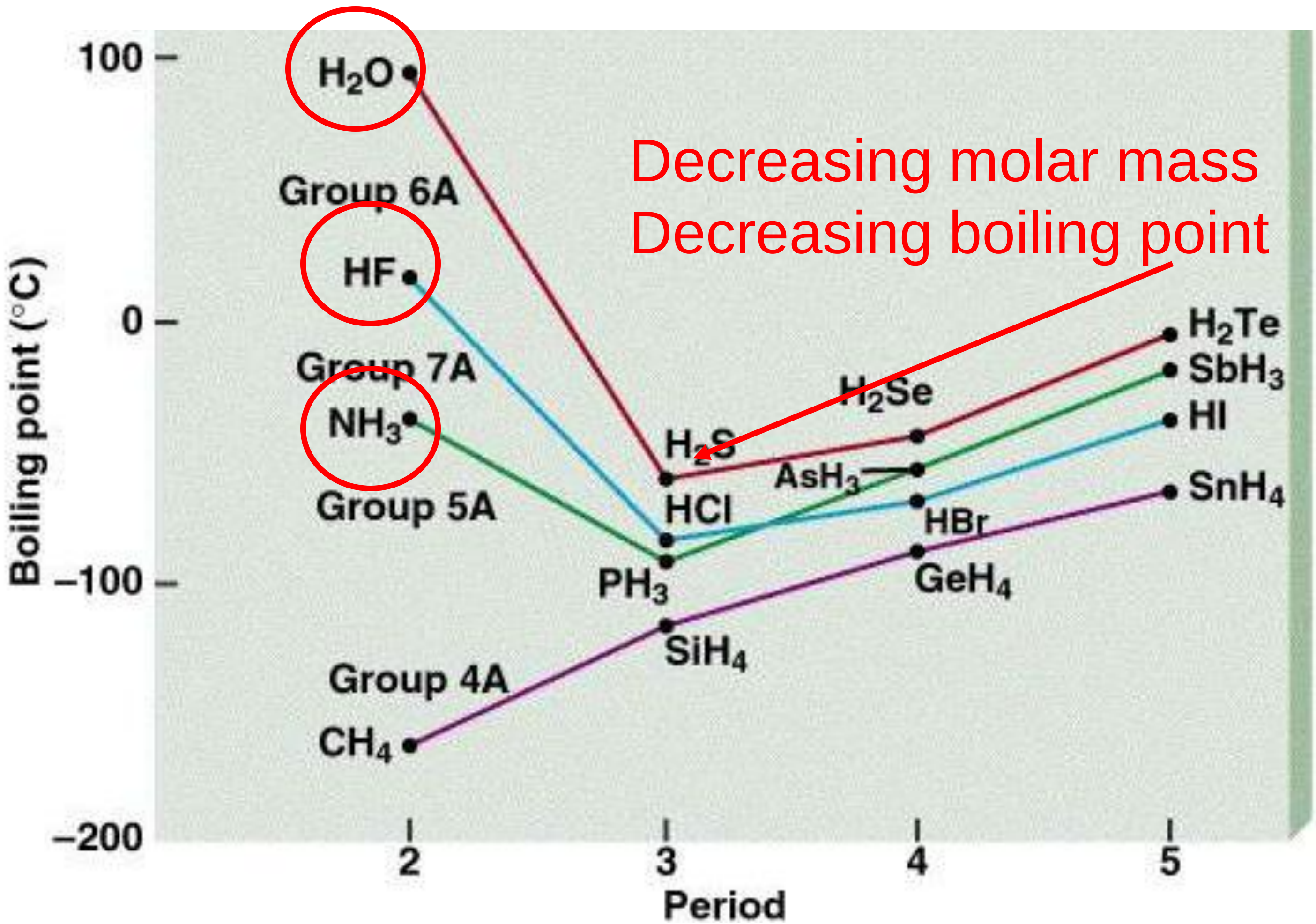
Maximum Density
4^0C

Density of Water





Why is the hydrogen bond considered a "special" dipole-dipole interaction?

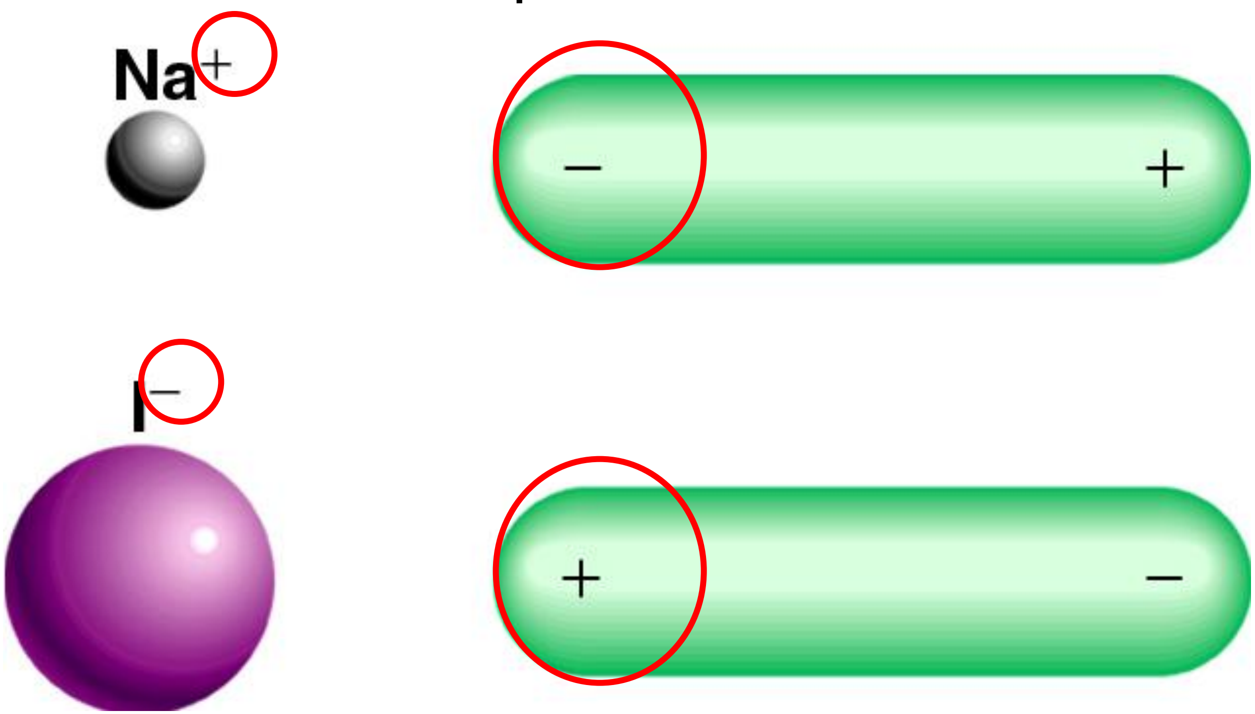


Types of Intermolecular Forces

2. Ion-Dipole Forces

Attractive forces between an **ion** and a **polar molecule.** The magnitude of the attraction increases as either the charge of the ion or the magnitude of the dipole increases.

Ion-Dipole Interaction

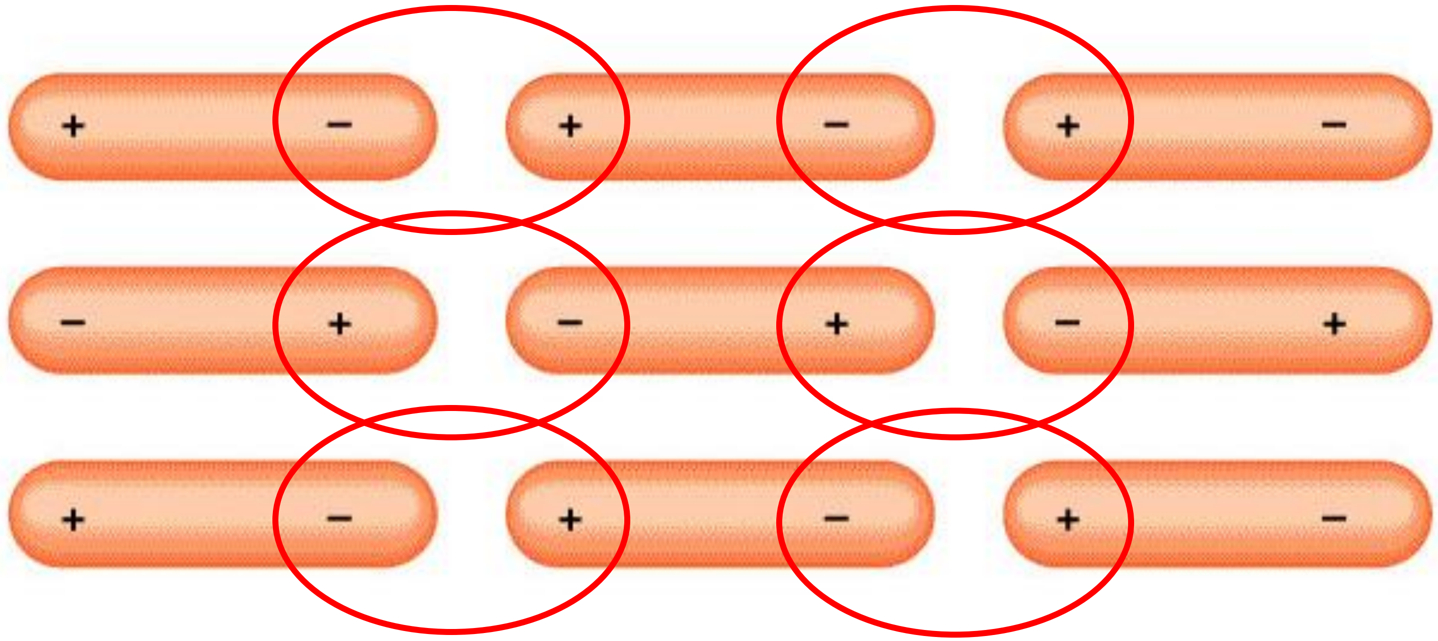


Types of Intermolecular Forces

3. Dipole-Dipole Forces

Attractive forces between **polar molecules**

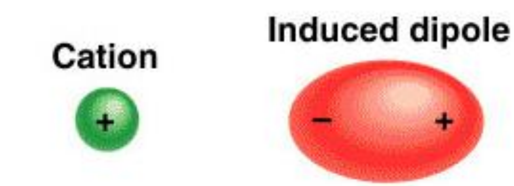
Orientation of Polar Molecules in a Solid



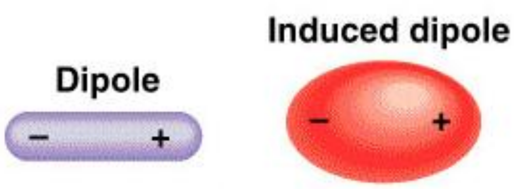
Types of Intermolecular Forces

4. Dispersion Forces – van der Walls forces/London forces (weakest)

Attractive forces that arise as a result of **temporary dipoles induced** in atoms or molecules



ion-induced dipole interaction



dipole-induced dipole interaction

Intermolecular Forces

4. Dispersion Forces Continued (see Animation)

Polarizability is the ease with which the electron distribution in the atom or molecule can be distorted.

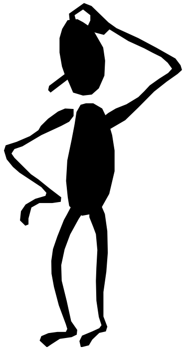
Polarizability increases with:

- greater number of electrons
- more diffuse electron cloud

Dispersion forces usually increase with molar mass.

TABLE 11.2 Melting Points of Similar Nonpolar Compounds

Compound	Melting Point (°C)
CH_4	−182.5
CF_4	−150.0
CCl_4	−23.0
CBr_4	90.0
CI_4	171.0



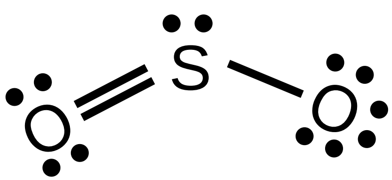
What type(s) of intermolecular forces exist between each of the following molecules?

HBr

HBr is a polar molecule: dipole-dipole forces. There are also dispersion forces between HBr molecules.

CH_4

CH_4 is nonpolar: dispersion forces.



SO_2

SO_2 is a polar molecule: dipole-dipole forces. There are also dispersion forces between SO_2 molecules.

Intermolecular Forces

Intermolecular forces are attractive forces **between** molecules.

Intramolecular forces hold atoms together in a molecule.

Intermolecular vs Intramolecular

- 41 kJ to vaporize 1 mole of water (**inter**)
- 930 kJ to break all O-H bonds in 1 mole of water (**intra**)

Generally, **inter**molecular forces are much weaker than **intra**molecular forces.

"Measure" of intermolecular force

boiling point

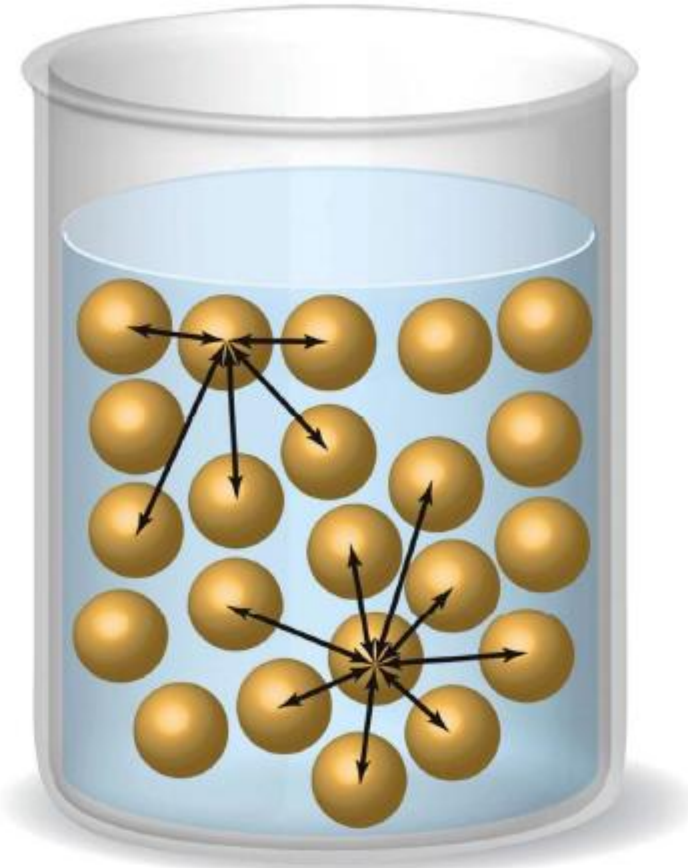
melting point

ΔH_{vap}

ΔH_{fus}

ΔH_{sub}

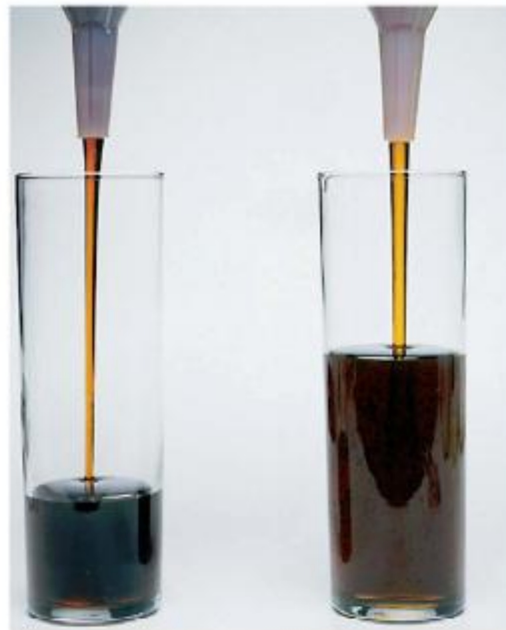
Intermolecular Forces Affect Many Physical Properties



The strength of the attractions between particles can greatly affect the properties of a substance or solution.

Viscosity

- **Resistance of a liquid to flow is called viscosity.**
- **It is related to the ease with which molecules can move past each other.**
- **Viscosity:**
  - **Increases with stronger intermolecular forces**
  - **Increases with the size of the molecules**
  - **Decreases with increasing temperature.**



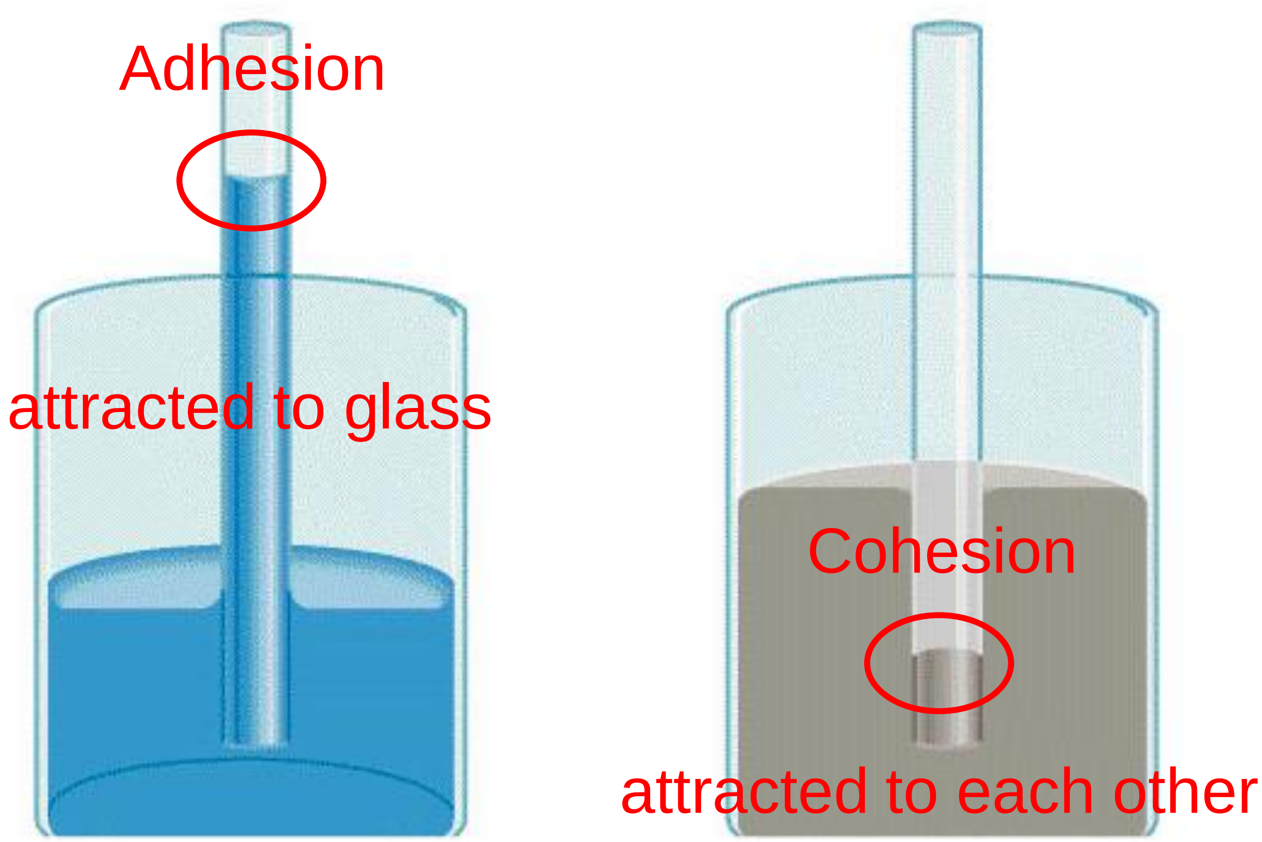
Two methods of measuring viscosity:

1. **Timing the flow of a liquid through an opening.**
2. **A disk or drum type viscometer**

Properties of Liquids

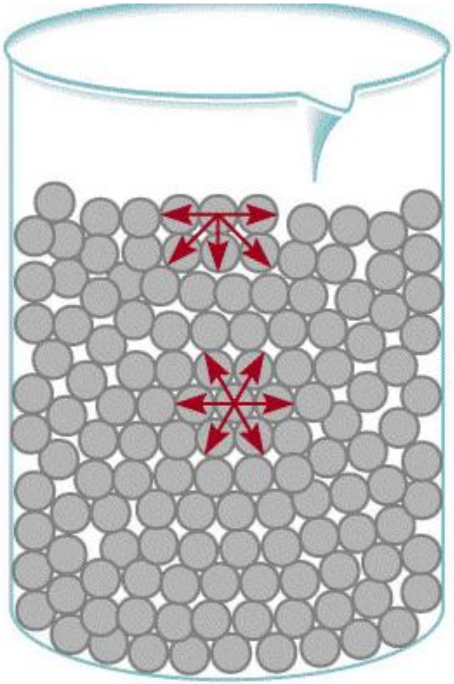
Cohesion is the intermolecular attraction between like molecules

Adhesion is an attraction between unlike molecules



Properties of Liquids

Surface tension is the amount of energy required to stretch or increase the surface of a liquid by a unit area. It results from the net inward force experienced by the molecules on the surface of a liquid.

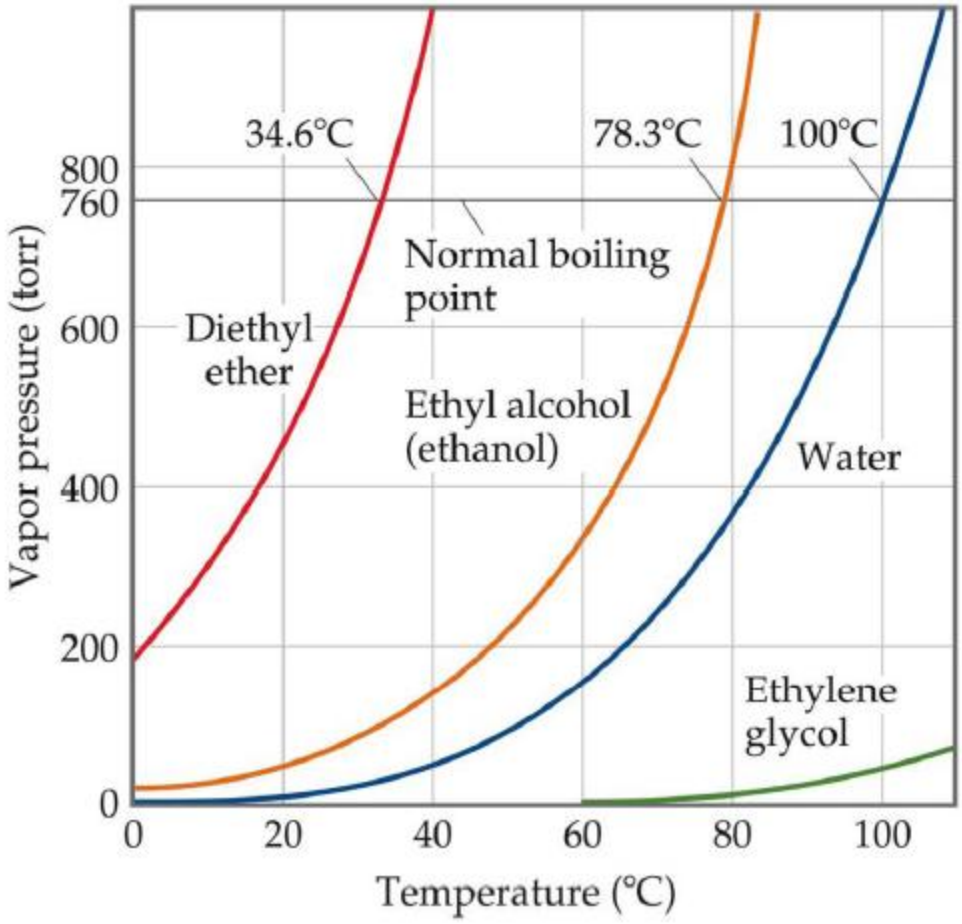


Strong intermolecular forces

High surface tension

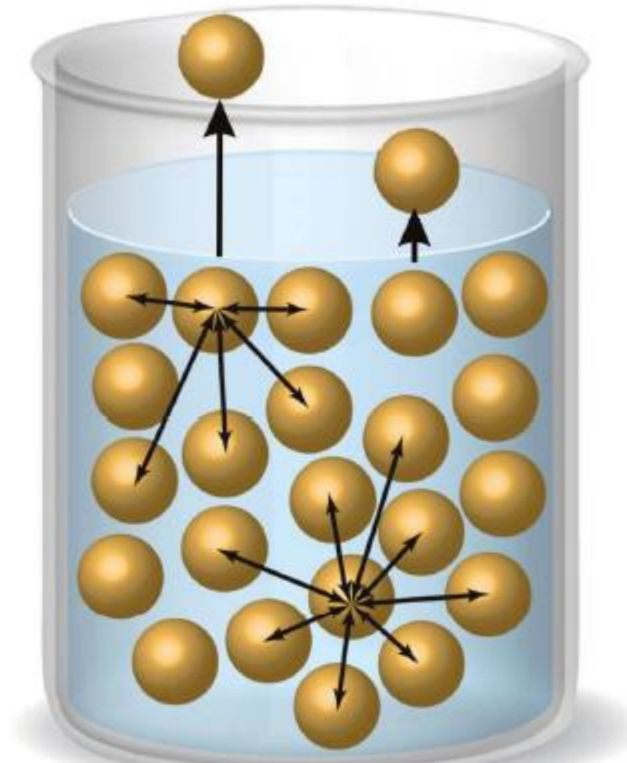
Vapor Pressure

- Vapor pressure increases with temperature.
- When the vapor pressure of a liquid equals the atmospheric pressure, the liquid boils.
- The normal boiling point of a liquid is the temperature at which its vapor pressure is 760 torr.



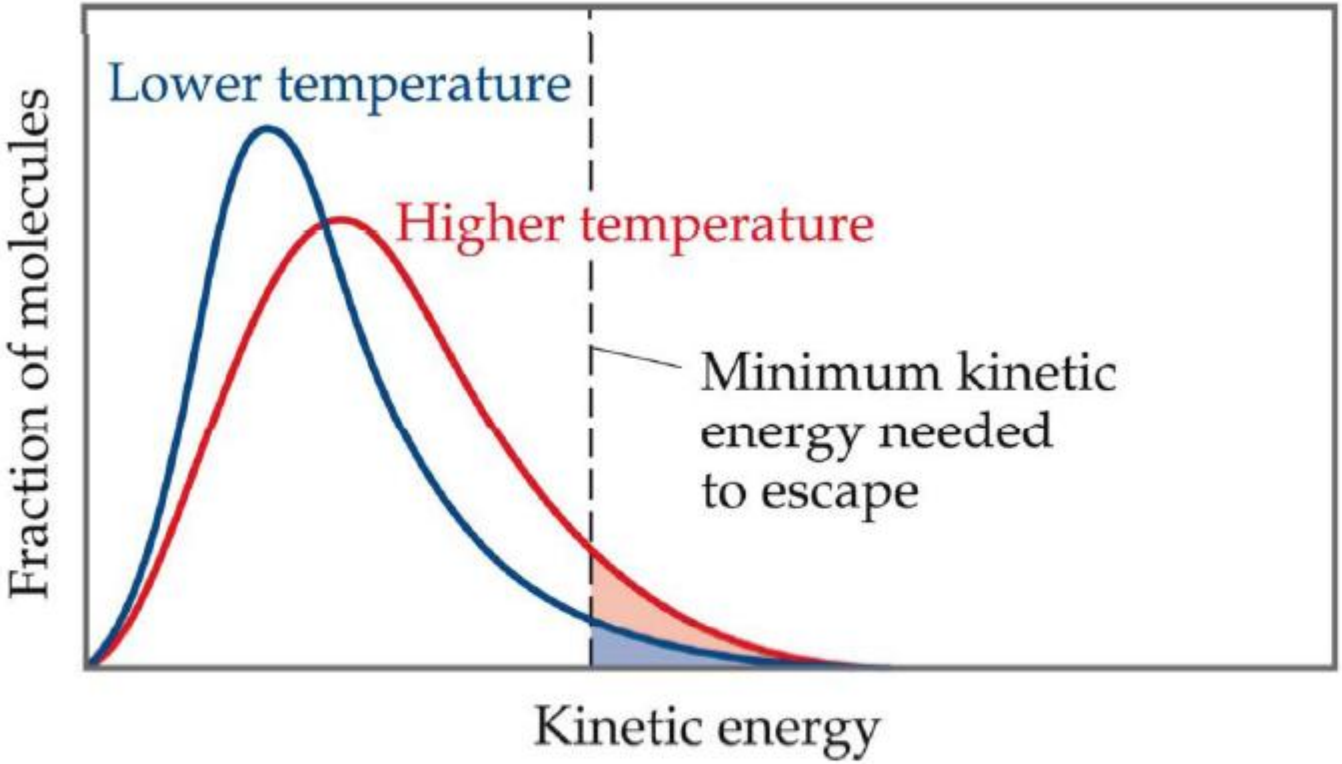
Vapor Pressure

Due to both temperature effects and energy transfers from collisions, molecules on the surface of a liquid are able to gain sufficient kinetic energy to escape into the atmosphere



Vapor Pressure

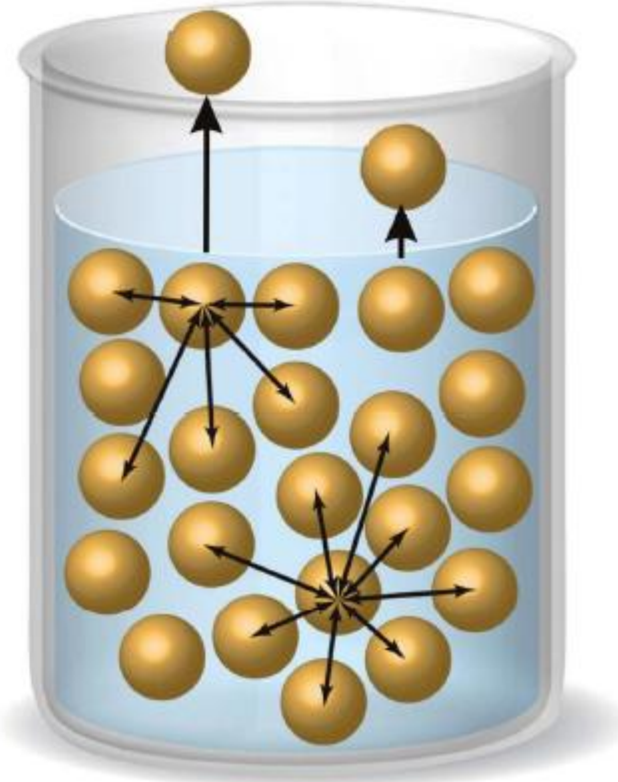
- At any temperature, some molecules in a liquid have enough energy to escape.
- As the temperature rises, the fraction of molecules that have enough energy to escape increases.



Vapor Pressure

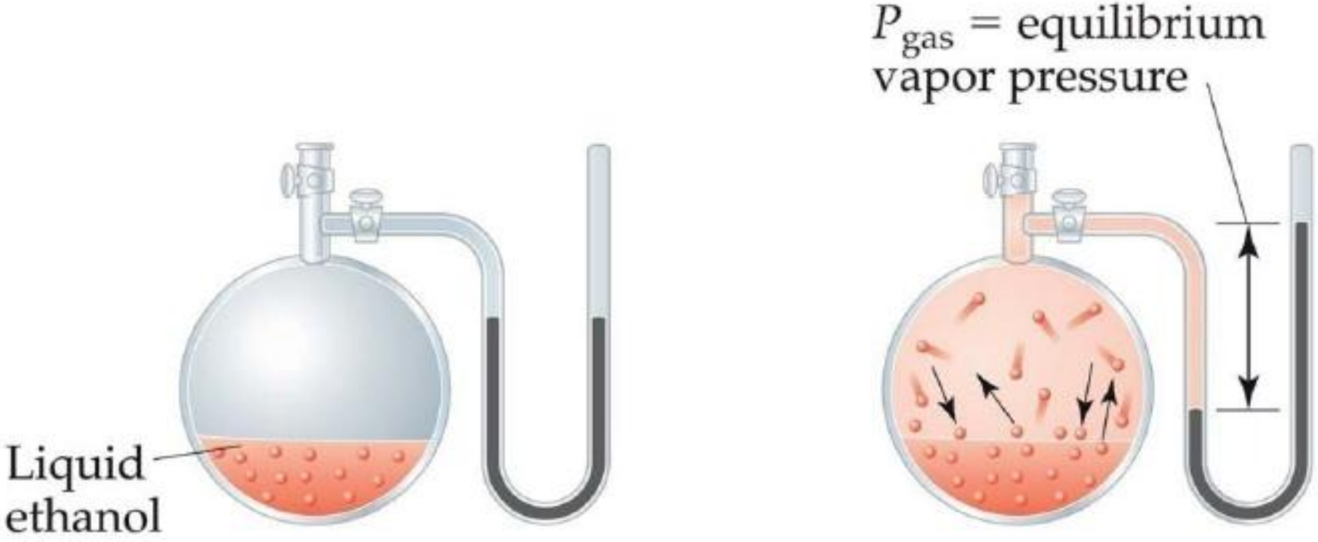
If the container is open to the atmosphere, the molecules simply escape. This process is called evaporation.

As molecules escape from the surface, they take energy with them resulting in a cooling effect on the liquid.



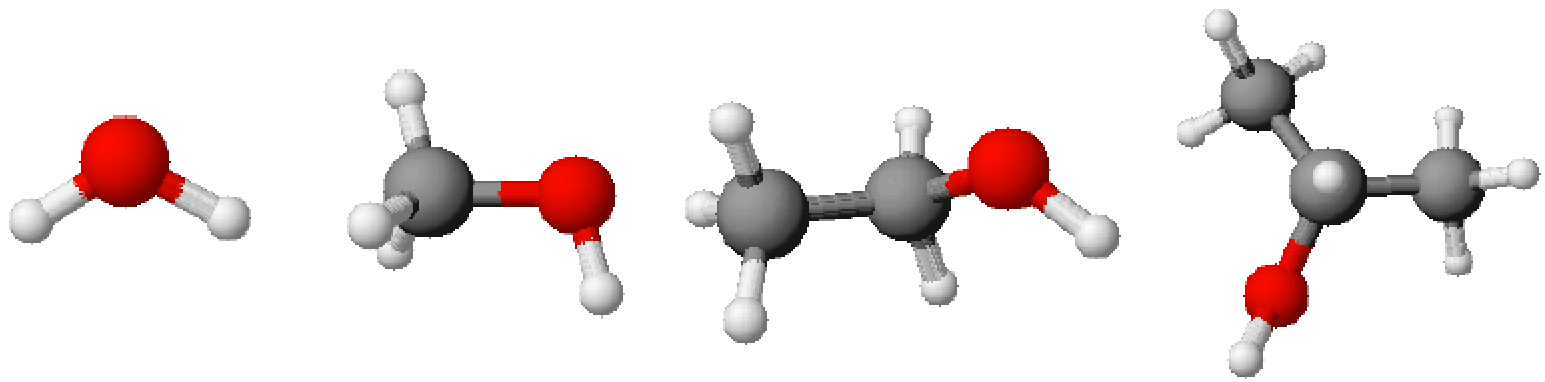
Vapor Pressure

If the container is closed to the atmosphere, as more molecules escape the liquid, the pressure they exert increases.



Which Will Evaporate First?

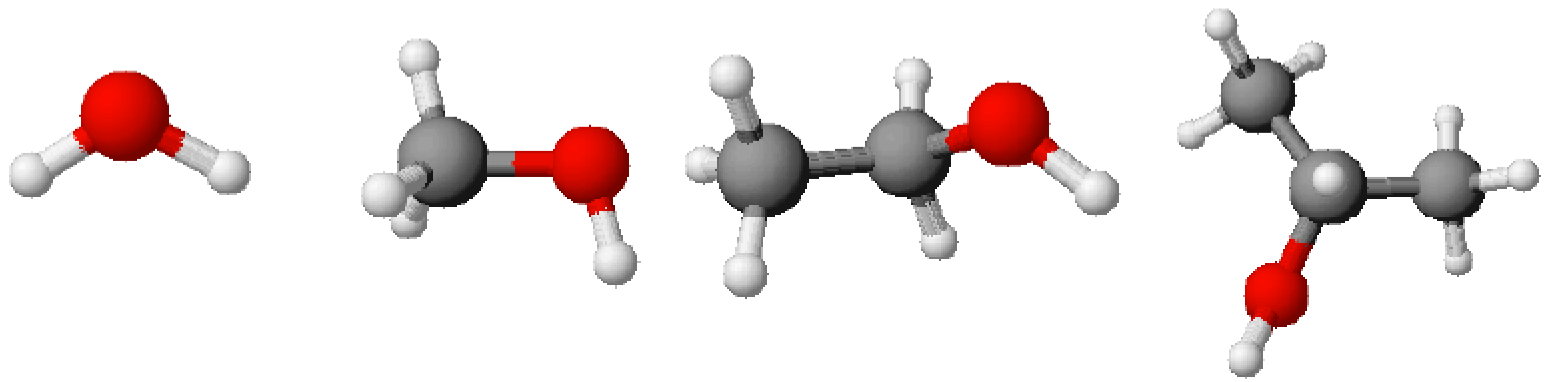
What factors affect evaporation?



Water methanol ethanol 2-propanol

Which Will Evaporate First?

What factors affect evaporation?



Water **methanol** **ethanol** **2-propanol**

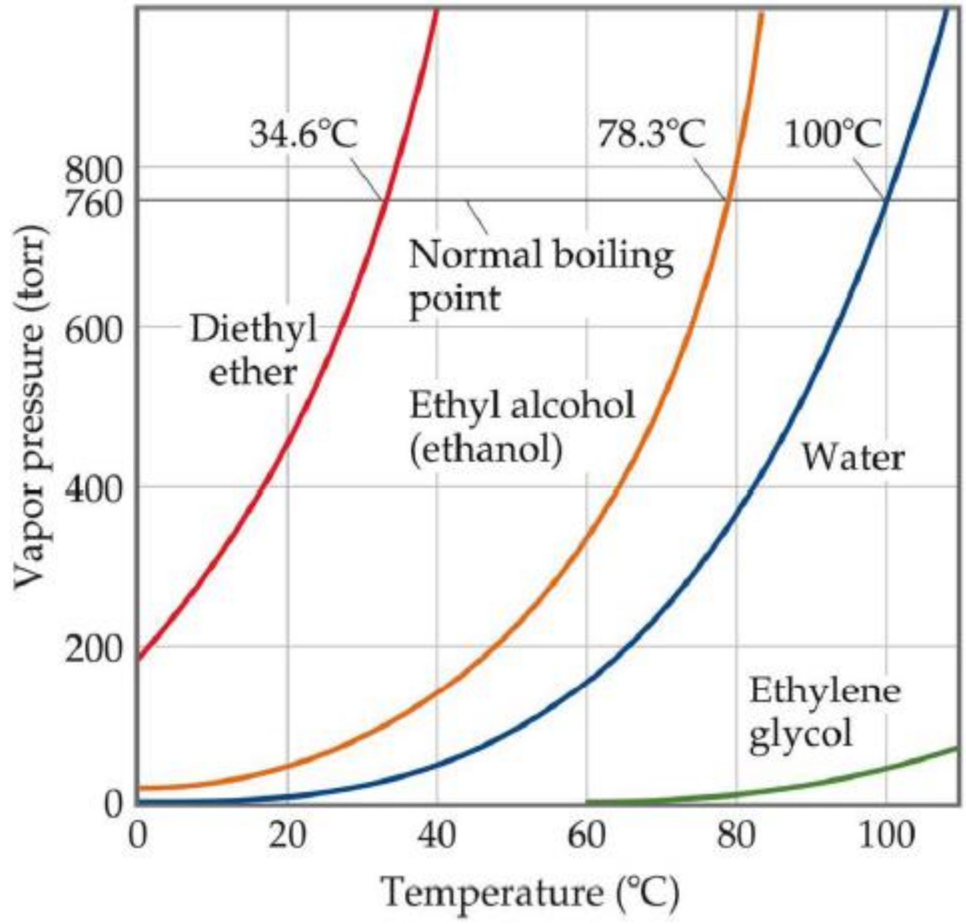
Effect of molecular weight:

$H_2O = 18$ $CH_3OH = 32$ $C_2H_5OH = 46$ $C_3H_8OH = 60$

Effect of polarity

Vapor Pressure

- Vapor pressure increases with temperature.
- When the vapor pressure of a liquid equals the atmospheric pressure, the liquid boils.
- The normal boiling point of a liquid is the temperature at which its vapor pressure is 760 torr.



Vapor pressure of water versus temperature

Atmospheric pressure at various altitudes

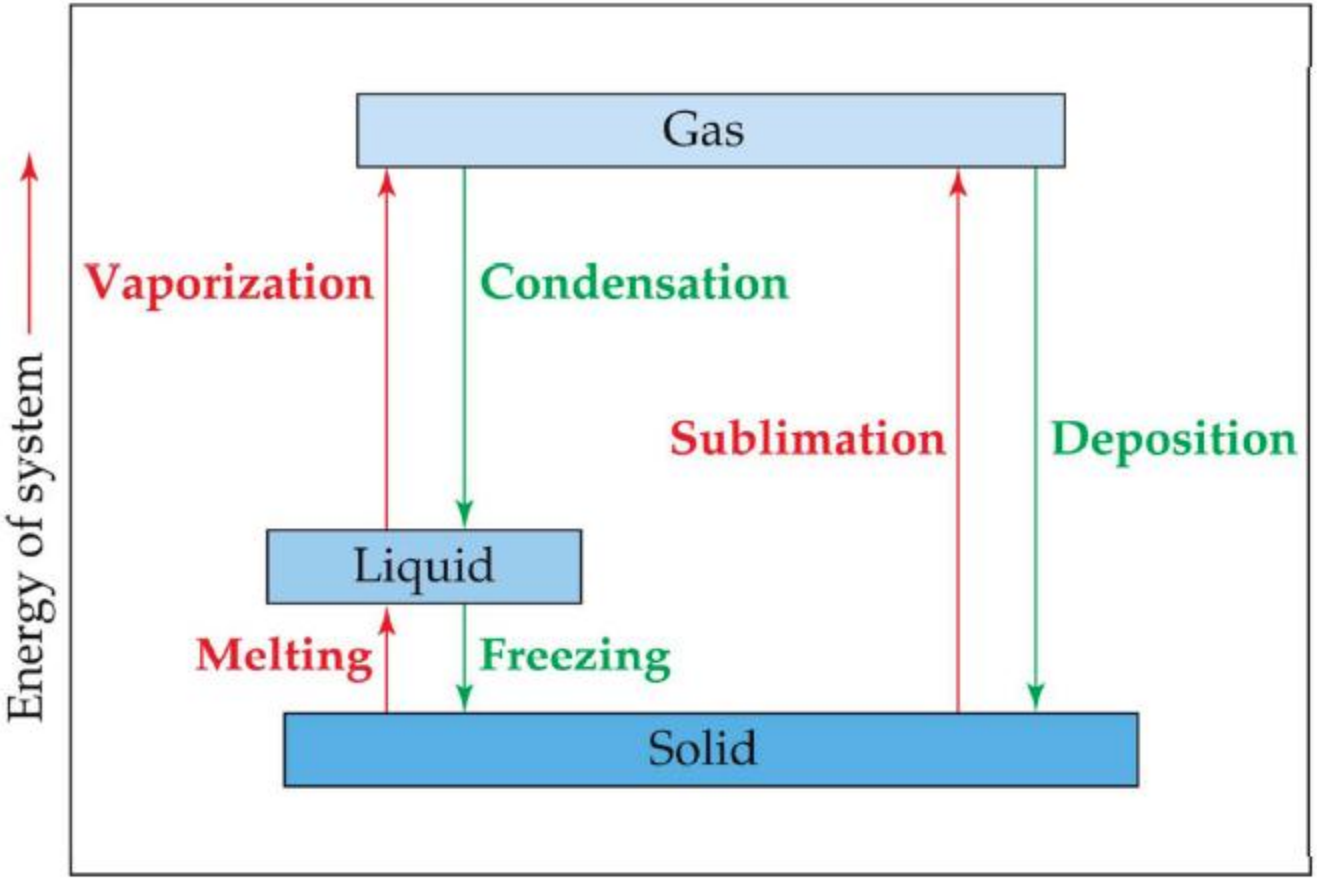
Vapor pressure of water at various temperatures

Temp. °C.	0.0° mm of Hg	0.2° mm of Hg	0.4° mm of Hg	0.6° mm of Hg	0.8° mm of Hg
80	355.40	358.28	361.19	364.11	367.06
81	370.03	373.01	376.02	379.05	382.09
82	385.16	388.25	391.36	394.49	397.64
83	400.81	404.00	407.22	410.45	413.71
84	416.99	420.29	423.61	426.95	430.32
85	433.71	437.12	440.55	444.01	447.49
86	450.99	454.51	458.06	461.63	465.22
87	468.84	472.48	476.14	479.83	483.54
88	487.28	491.04	494.82	498.63	502.46
89	506.32	510.20	514.11	518.04	521.99
90	525.97	529.98	534.01	538.07	542.15
91	546.26	550.40	554.56	558.75	562.96
92	567.20	571.47	575.76	580.08	584.43
93	588.80	593.20	597.63	602.09	606.57
94	611.08	615.62	620.19	624.79	629.41
95	634.06	638.74	643.45	648.19	652.96
96	657.75	662.58	667.43	672.32	677.23
97	682.18	687.15	692.15	697.19	702.25
98	707.35	712.47	717.63	722.81	728.03
99	733.28	738.56	743.87	749.22	754.59
100	760.00	765.44	770.91	776.42	781.95

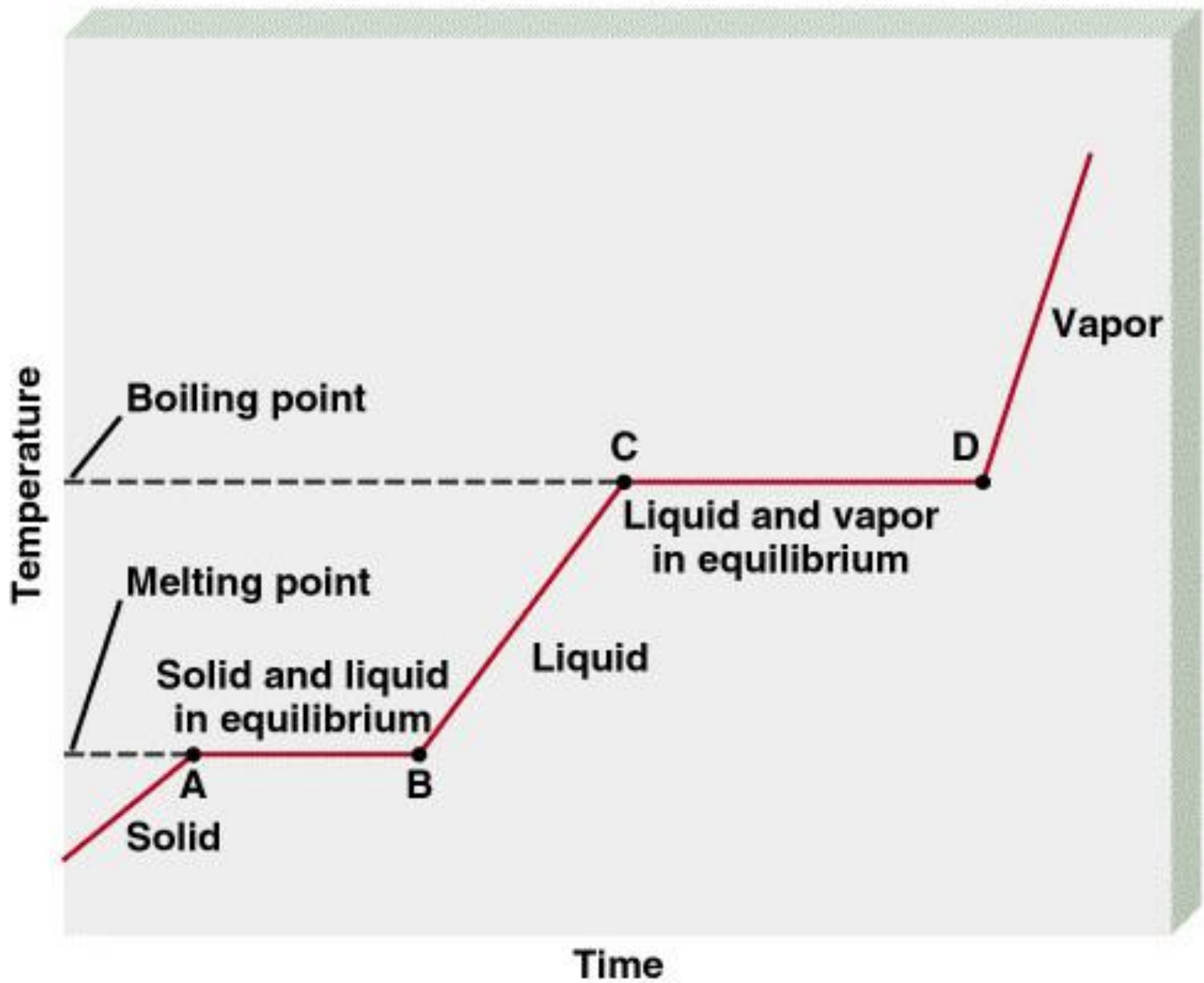
Atmospheric pressure at various altitudes

	Feet	Meters	Atm	Mm Hg
Sea level ➔	0	0	1.00	760
	328	100	0.99	752
	500	150	0.98	747
	656	200	0.97	743
	1000	300	0.96	734
	1312	400	0.95	725
	1500	450	0.94	719
	2000	600	0.93	706
Tucson ➔	2500	750	0.91	694
	3000	900	0.89	681
	3500	1070	0.88	668
	4000	1220	0.86	655
	4500	1370	0.85	645
Denver ➔	5000	1520	0.83	633
	5500	1680	0.81	620
	6000	1830	0.80	610
	6500	1980	0.78	597
	7000	2130	0.77	587
	7500	2290	0.76	577
	8000	2440	0.74	564
	8500	2590	0.73	554

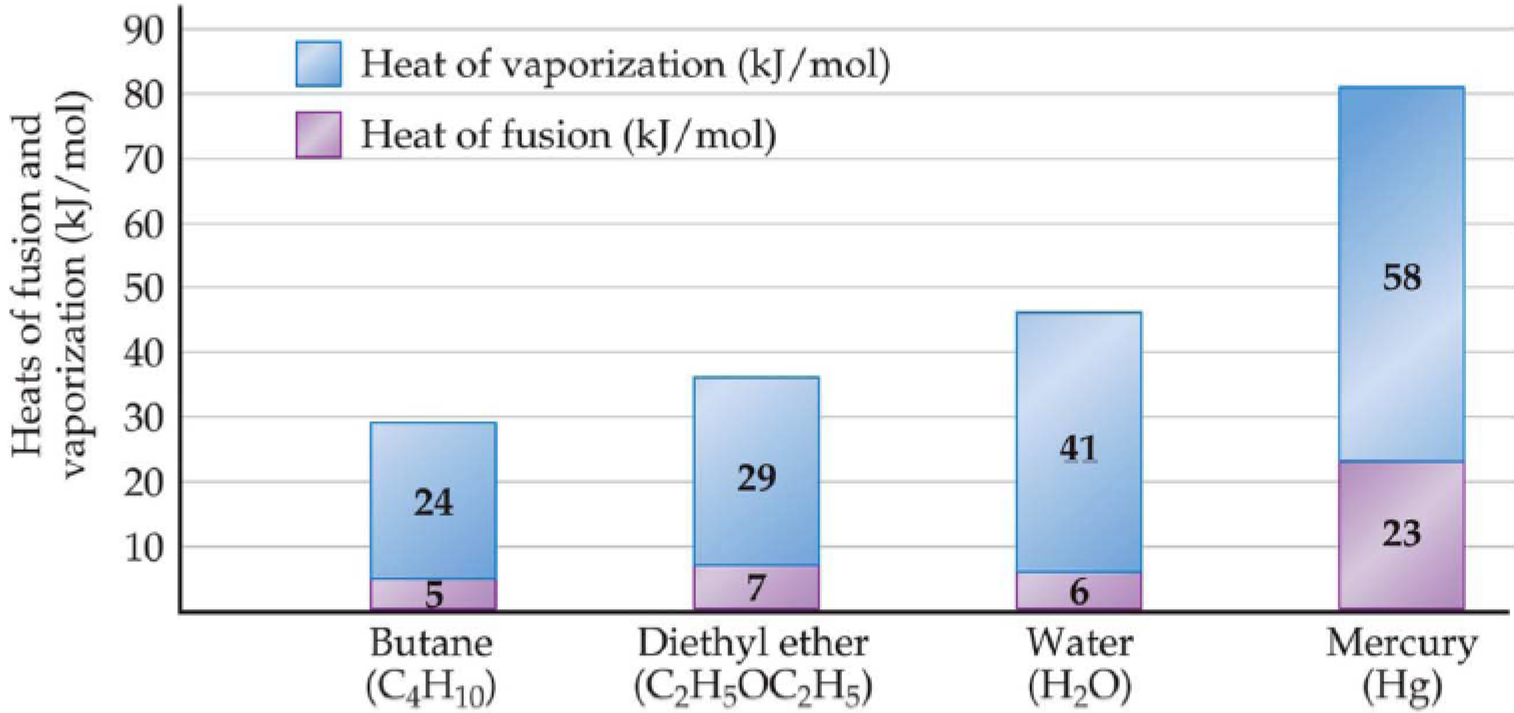
Phase Changes



Heating Curve



Energy Changes Associated with Changes of State



- Heat of Fusion: Energy required to change a solid at its melting point to a liquid.

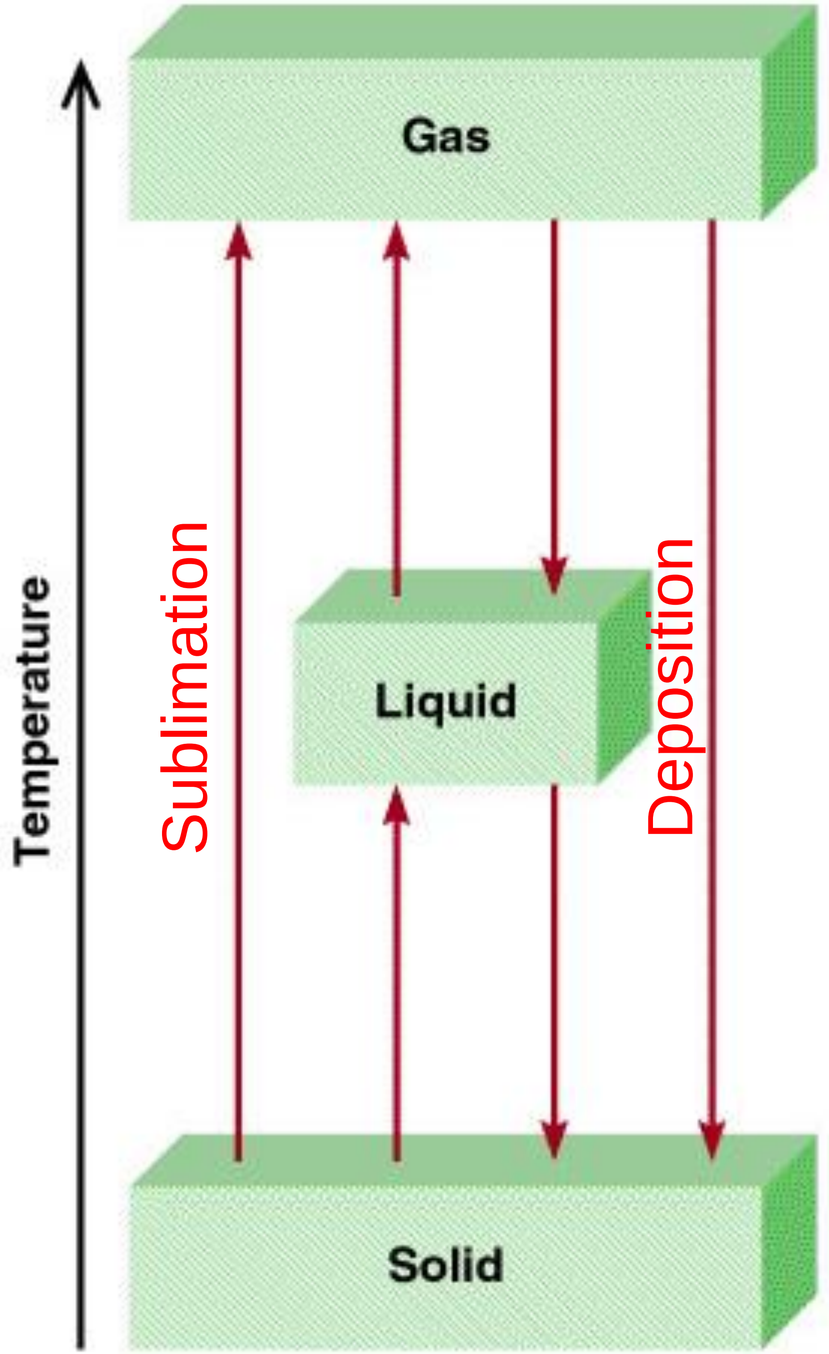
Phase Changes

$$H_2O\ (s) \rightleftharpoons H_2O\ (g)$$

Molar heat of sublimation (ΔH_{sub}) is the energy required to sublime 1 mole of a solid.

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

(Hess's Law)



Molar heat of fusion (ΔH_{fus}) is the energy required to melt 1 mole of a solid substance.

TABLE 11.8 Molar Heats of Fusion for Selected Substances

Substance	Melting Point* (°C)	ΔH_{fus} (kJ/mol)
Argon (Ar)	−190	1.3
Benzene (C_6H_6)	5.5	10.9
Ethanol (C_2H_5OH)	−117.3	7.61
Diethyl ether ($C_2H_5OC_2H_5$)	−116.2	6.90
Mercury (Hg)	−39	23.4
Methane (CH_4)	−183	0.84
Water (H_2O)	0	6.01

* Measured at 1 atm.

The ***boiling point*** is the temperature at which the (equilibrium) vapor pressure of a liquid is equal to the external pressure.

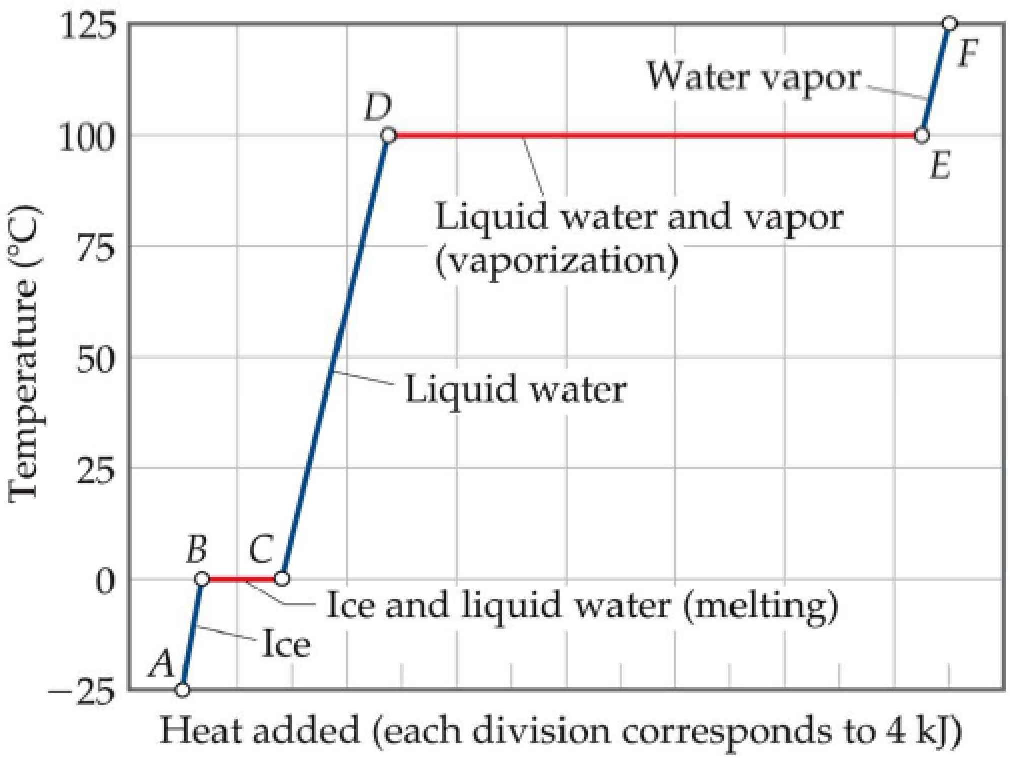
The ***normal boiling point*** is the temperature at which a liquid boils when the external pressure is 1 atm.

TABLE 11.6 Molar Heats of Vaporization for Selected Liquids

Substance	Boiling Point* (°C)	ΔH_{vap} (kJ/mol)
Argon (Ar)	−186	6.3
Benzene (C_6H_6)	80.1	31.0
Ethanol (C_2H_5OH)	78.3	39.3
Diethyl ether ($C_2H_5OC_2H_5$)	34.6	26.0
Mercury (Hg)	357	59.0
Methane (CH_4)	−164	9.2
Water (H_2O)	100	40.79

* Measured at 1 atm.

Energy Changes Associated with Changes of State



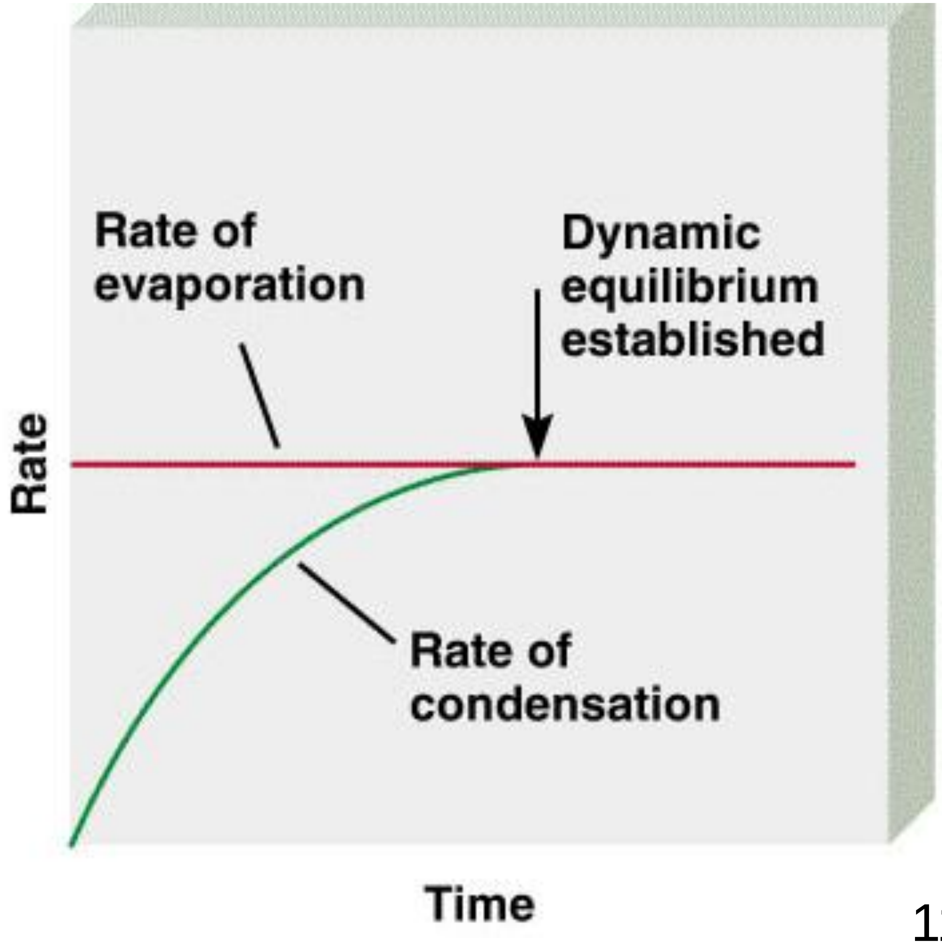
- Temperature remains constant at the melting and boiling points
  - Energy needed to break the intermolecular forces between the molecules.
  - Added kinetic energy for liquid or gaseous states.

The ***equilibrium vapor pressure*** is the vapor pressure measured when a dynamic equilibrium exists between condensation and evaporation

$$H_2O\ (l) \rightleftharpoons H_2O\ (g)$$

Dynamic Equilibrium

Rate of condensation = Rate of evaporation



Go To:
http://glencoe.com/sites/common_assets/advanced_placement/chemistry_chang10e/animations/vapor_pressure.swf

Molar heat of vaporization (ΔH_{vap}) is the energy required to vaporize 1 mole of a liquid.

Clausius-Clapeyron Equation

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

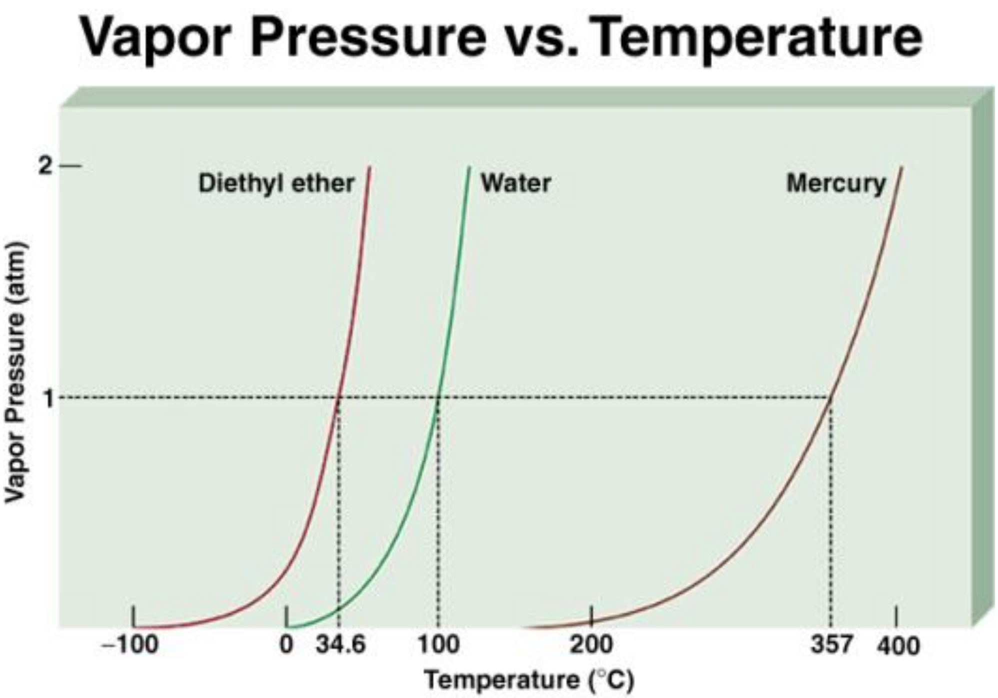
Is this equation on your AP reference tables?

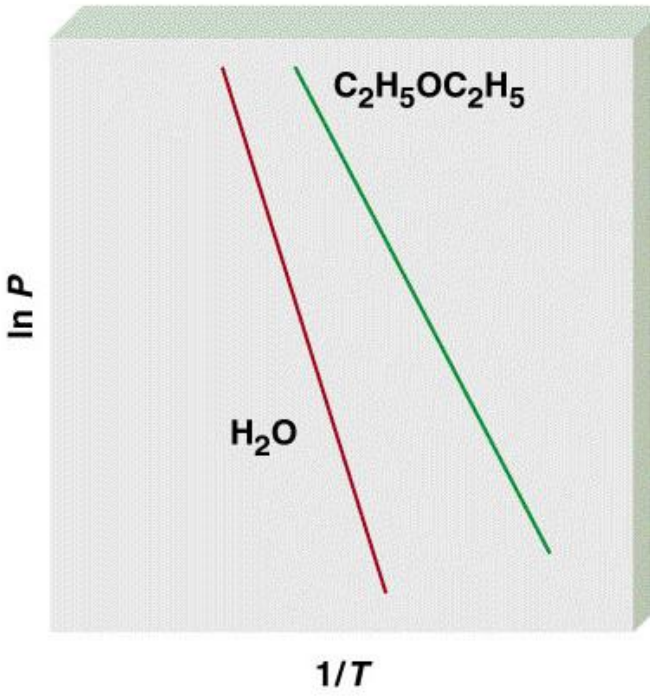
C = constant (depends on P & T)

P = (equilibrium) vapor pressure

T = temperature (K)

R = gas constant (8.314 J/K•mol)

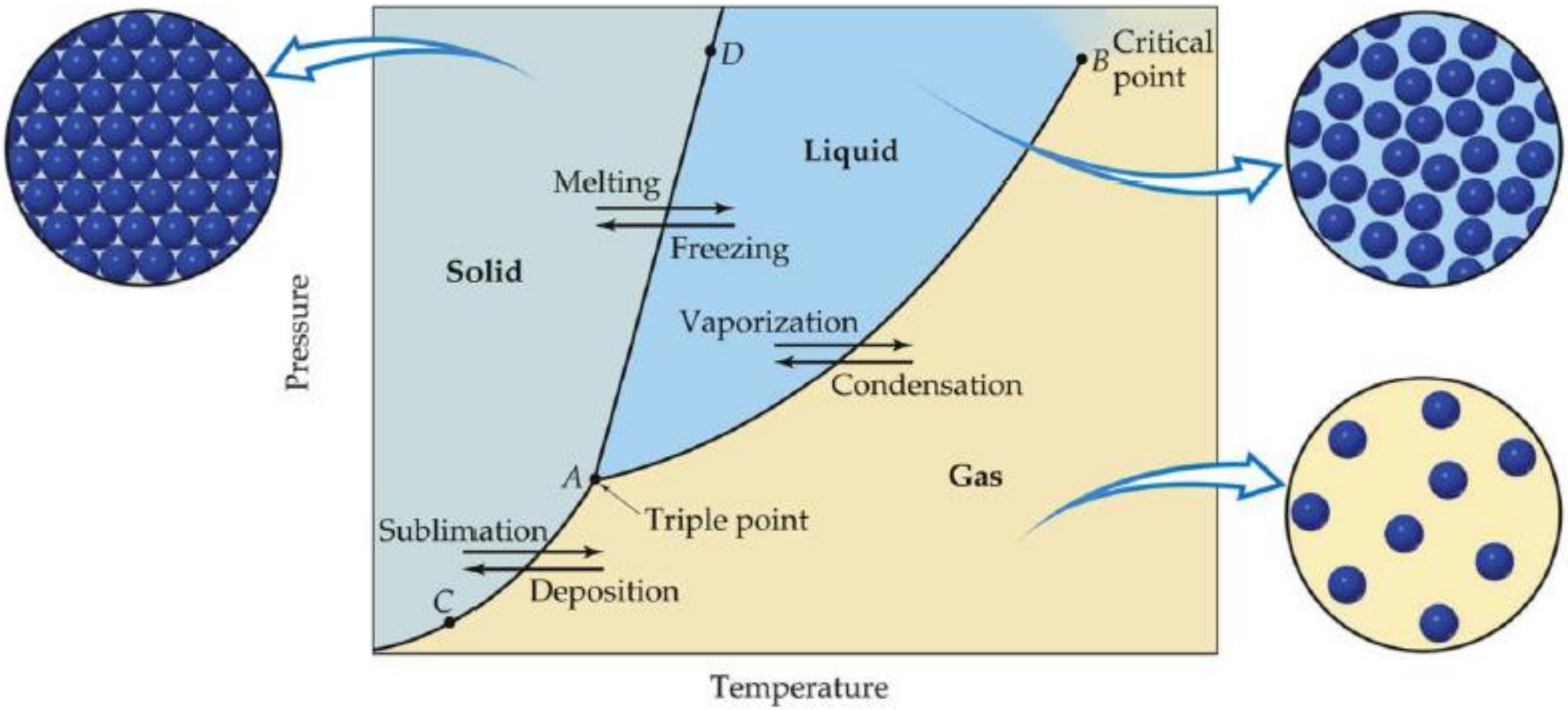




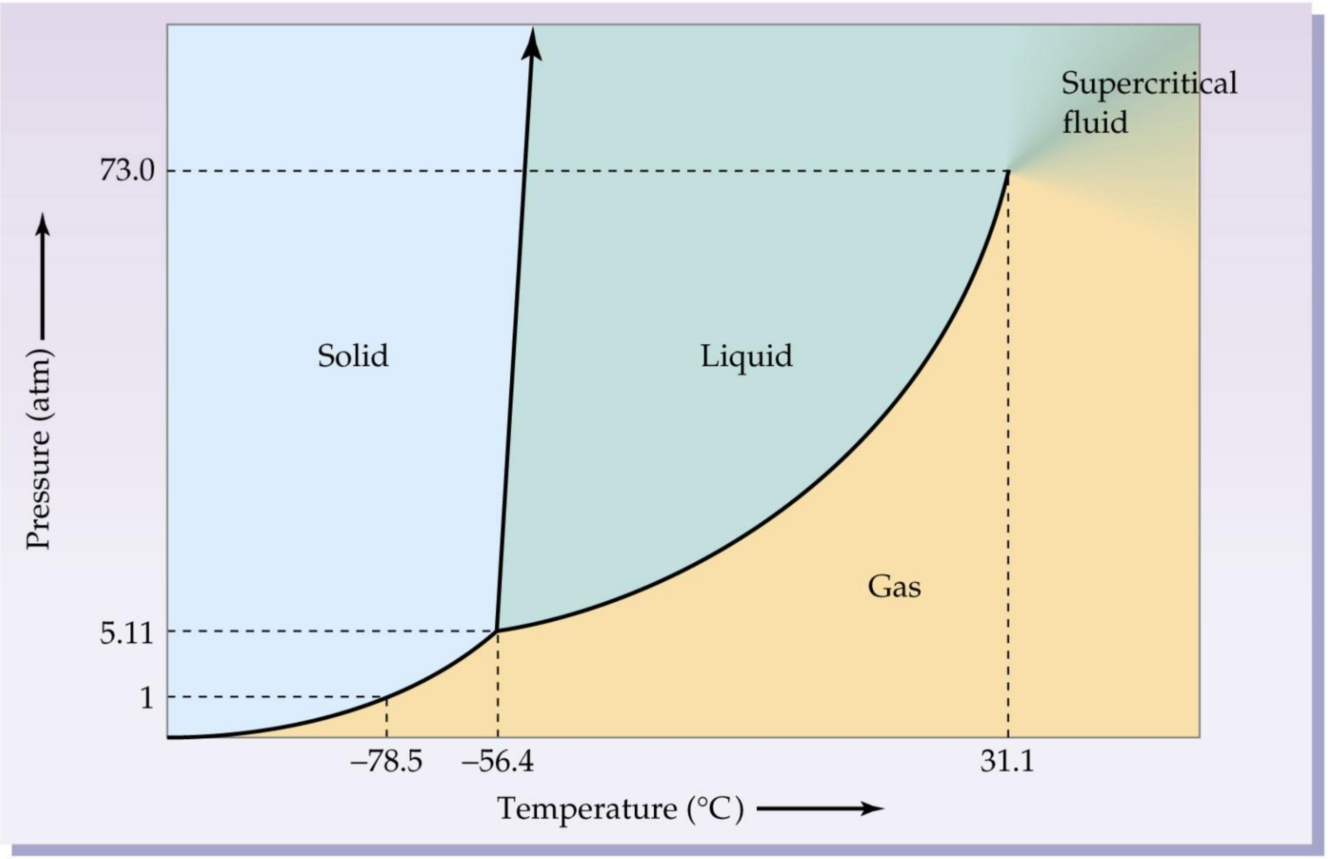
Phase Diagrams

Phase diagrams display the state of a substance at various pressures and temperatures and the places where equilibria exist between phases.

Each substance has its own unique phase diagram.



Where's Waldo?



Carbon Dioxide

Can you find…

The Triple Point?

Critical pressure?

Critical temperature?

Where fusion occurs?

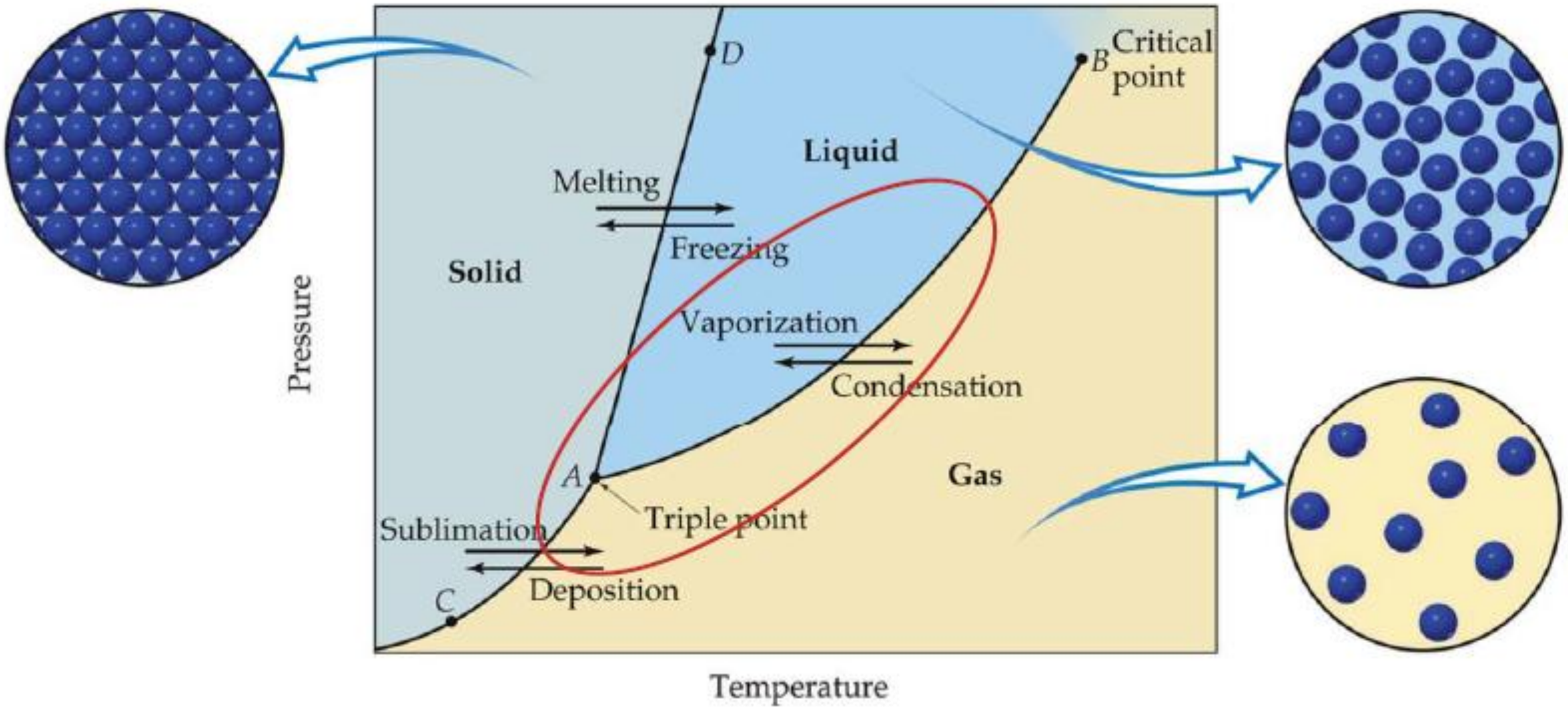
Where vaporization occurs?

Melting point (at 1 atm)?

Boiling point (at 6 atm)?

Phase Diagrams

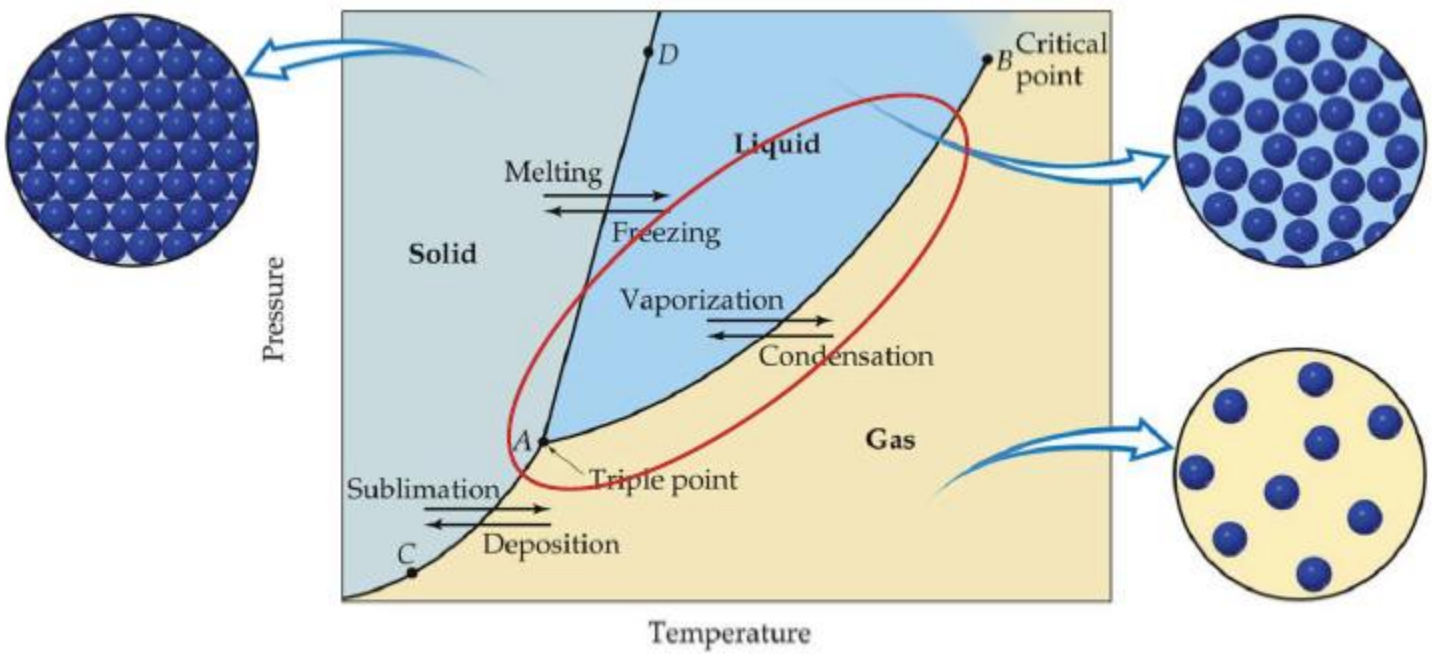
- The *AB* line is the liquid-vapor interface.
- It starts at the triple point (*A*), the point at which all three states are in equilibrium.



Phase Diagrams

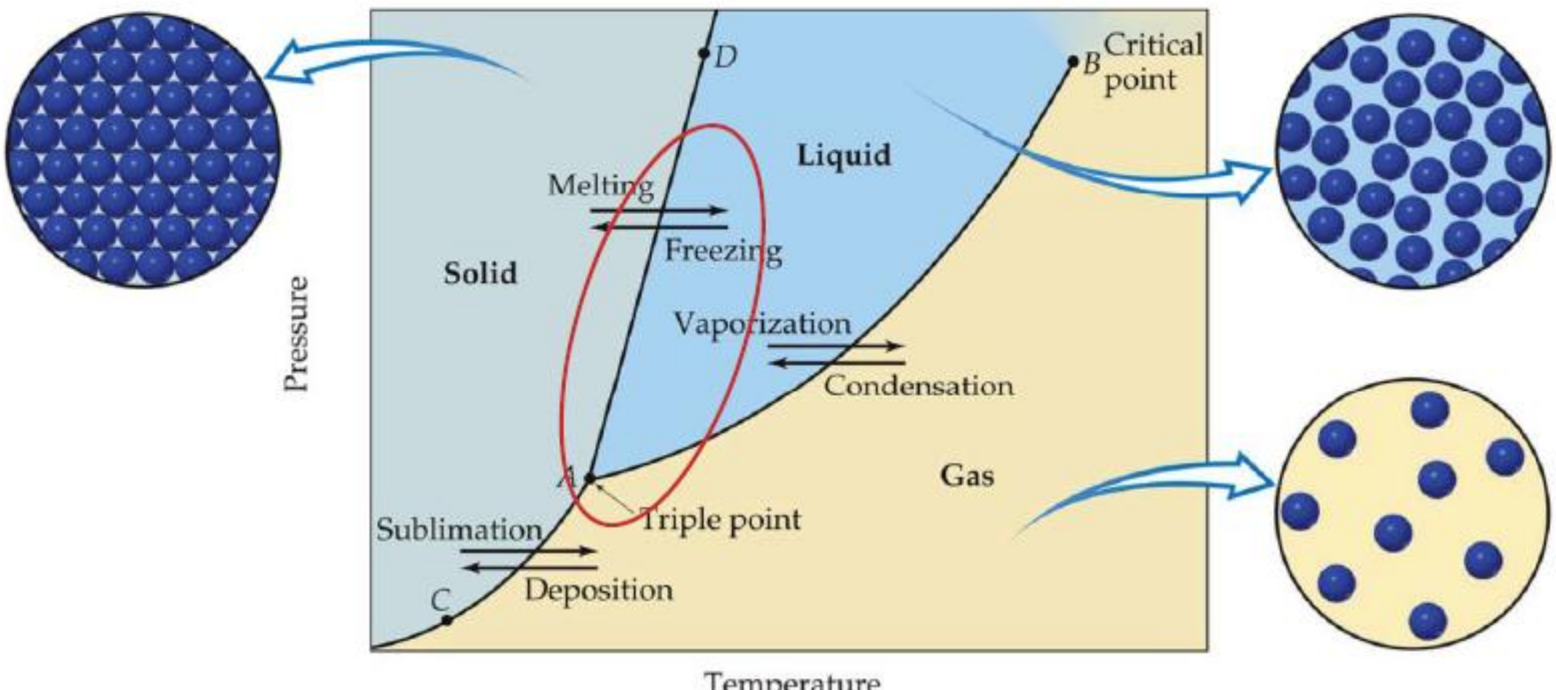
The critical point (*B*) is the highest temperature and pressure where the liquid form of the substance can exist.

Above the critical temperature and critical pressure the liquid and vapor are indistinguishable from each other.



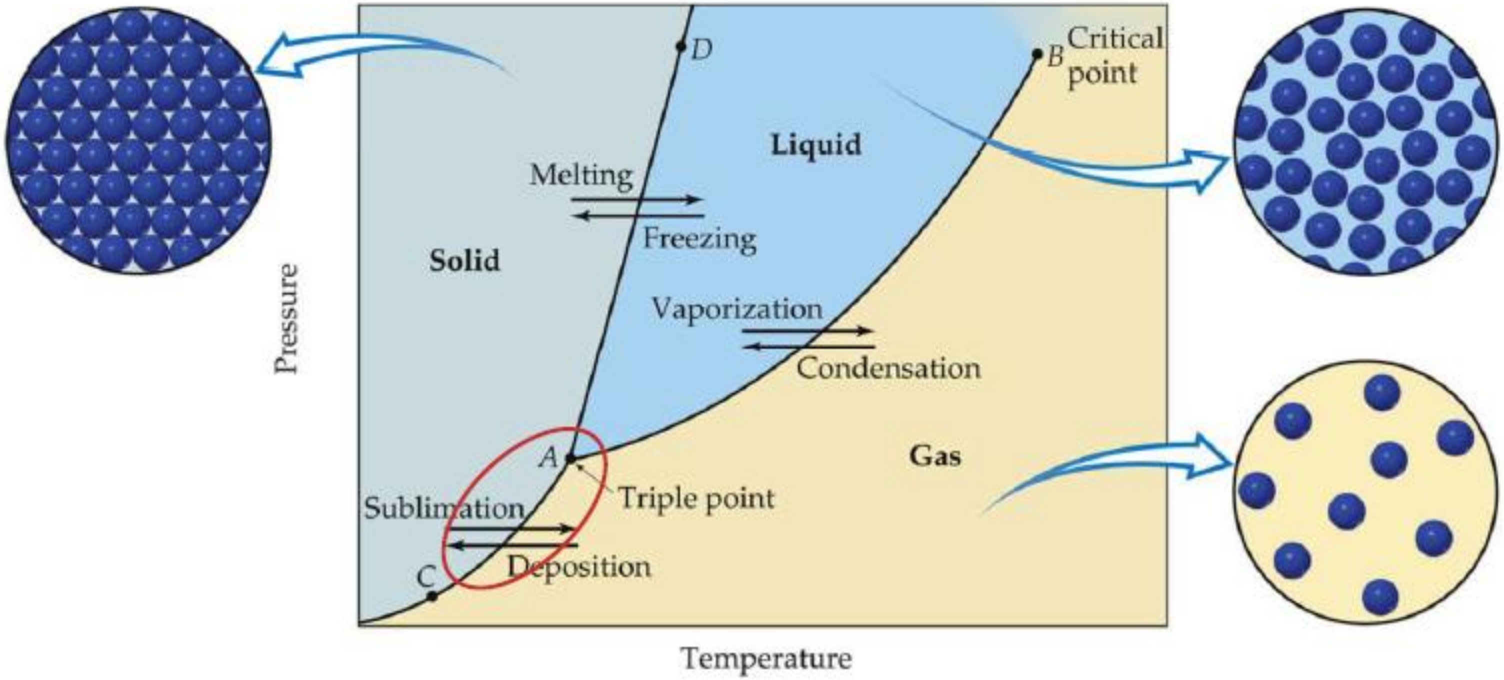
Phase Diagrams

- The *AD* line is the interface between liquid and solid.
- The melting point at each pressure can be found along this line.
- The substance represented in this phase diagram tends to decrease in volume on freezing, the melting point line slants to the right.

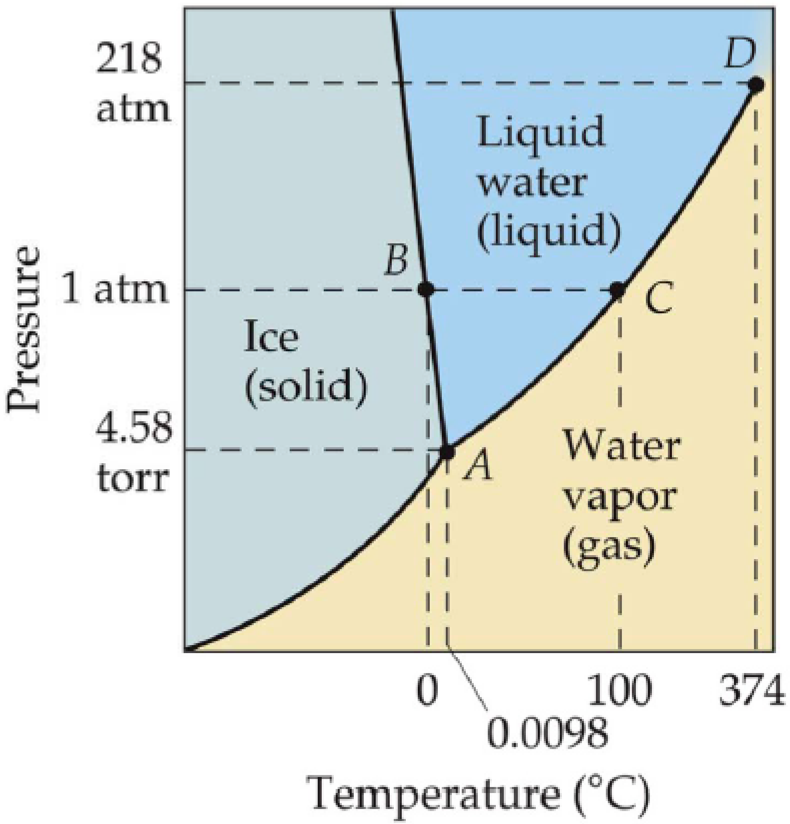


Phase Diagrams

- Below *A* the substance cannot exist in the liquid state.
- Along the *AC* line the solid and gas phases are in equilibrium.
- This is called the sublimation curve.

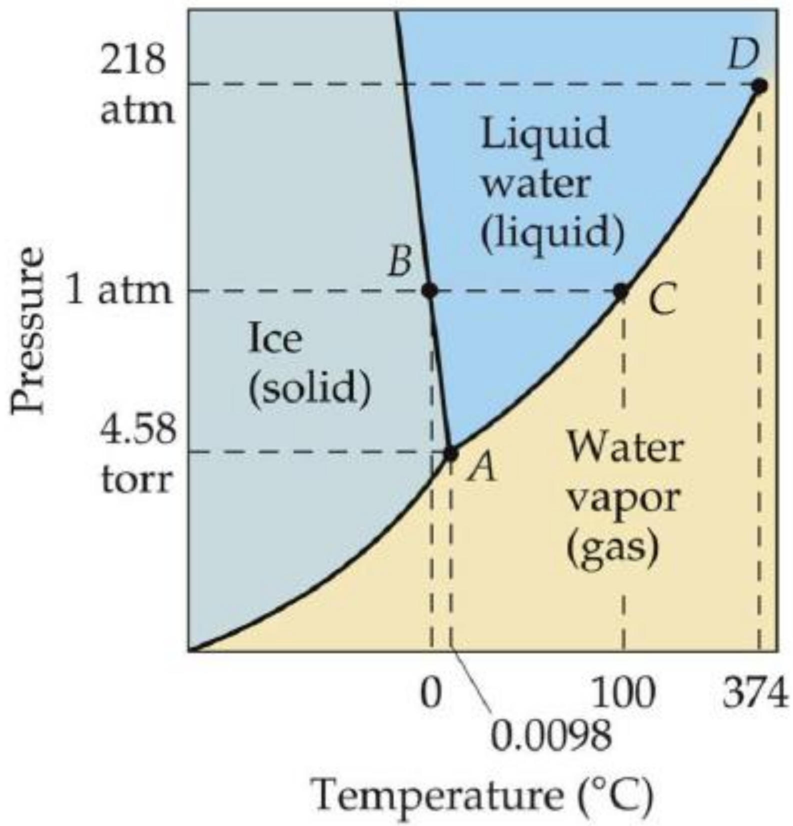


Phase Diagram of Water



- Note the high critical temperature and critical pressure:
  - These are due to the strong polar bonding between water molecules.
- Water expands on freezing, so the melting point line slants to the left.

Phase Diagram of Water

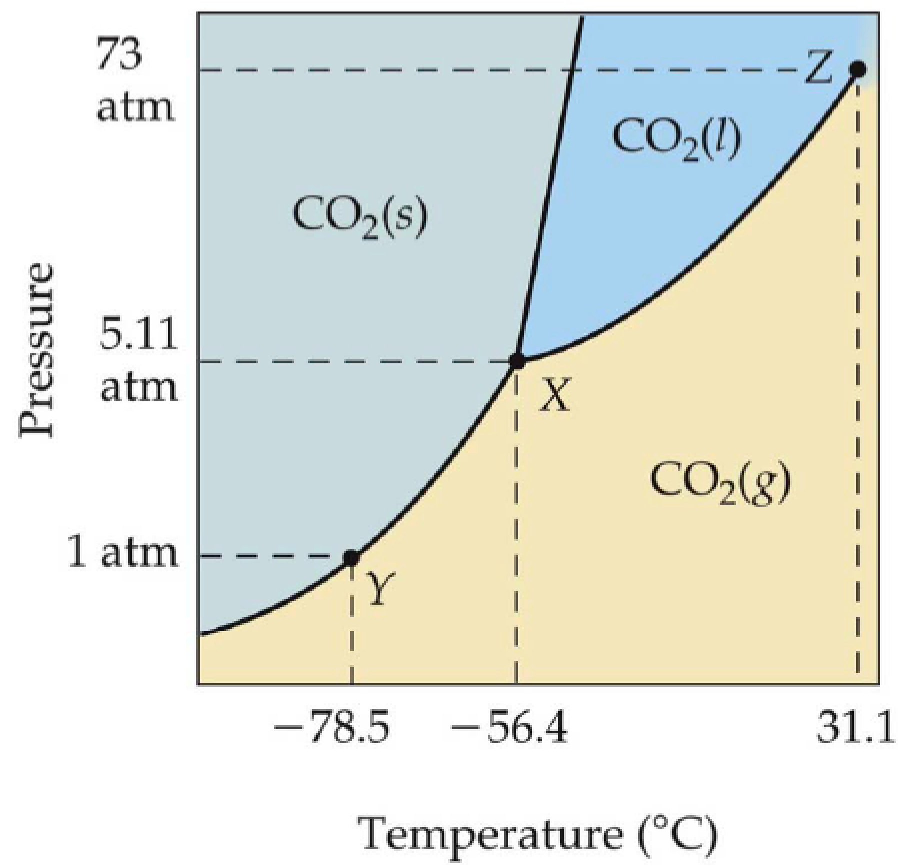


- The slope of the solid–liquid line is negative.
  - This means that as the pressure is increased at a temperature just below the melting point, water goes from a solid to a liquid.
  - This is why an ice skater can skate on ice.

Phase Diagram of Carbon Dioxide

Carbon dioxide cannot exist in the liquid state at pressures below 5.11 atm

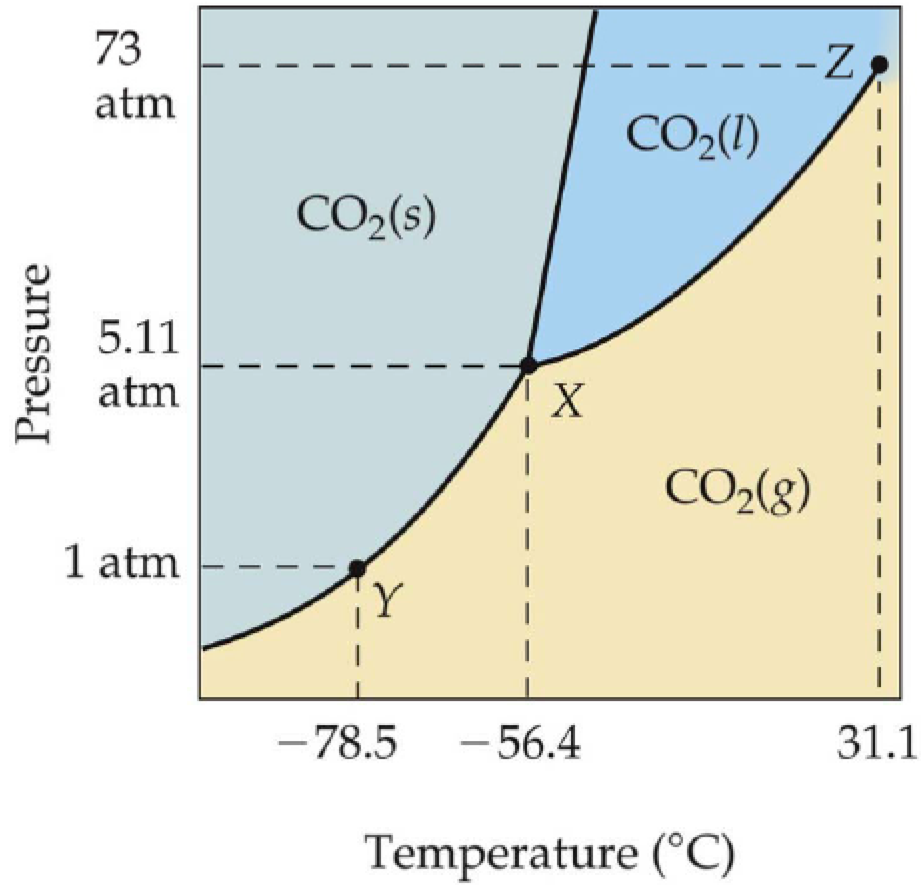
CO_2 sublimes at normal pressures.



Phase Diagram of Carbon Dioxide

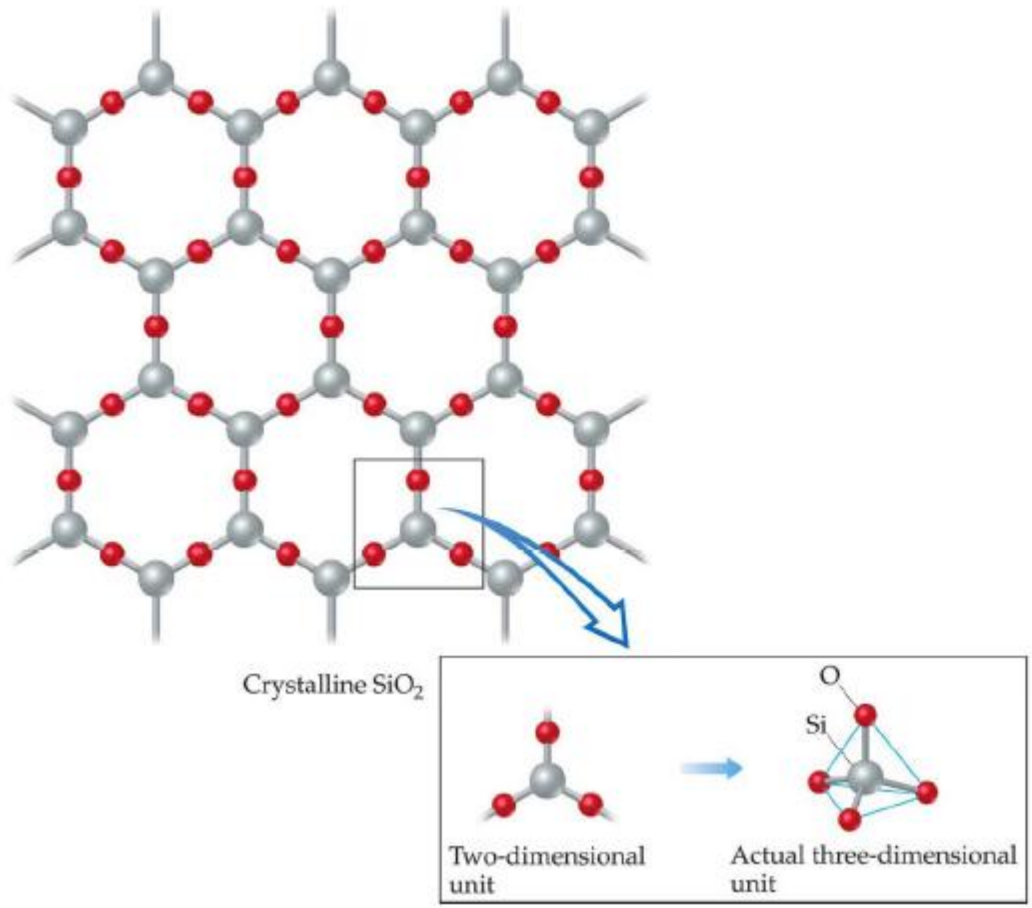
The low critical temperature and critical pressure for CO_2 make supercritical CO_2 a good solvent for extracting nonpolar substances (such as caffeine).

Supercritical CO_2 is being used for dry cleaning of clothing,



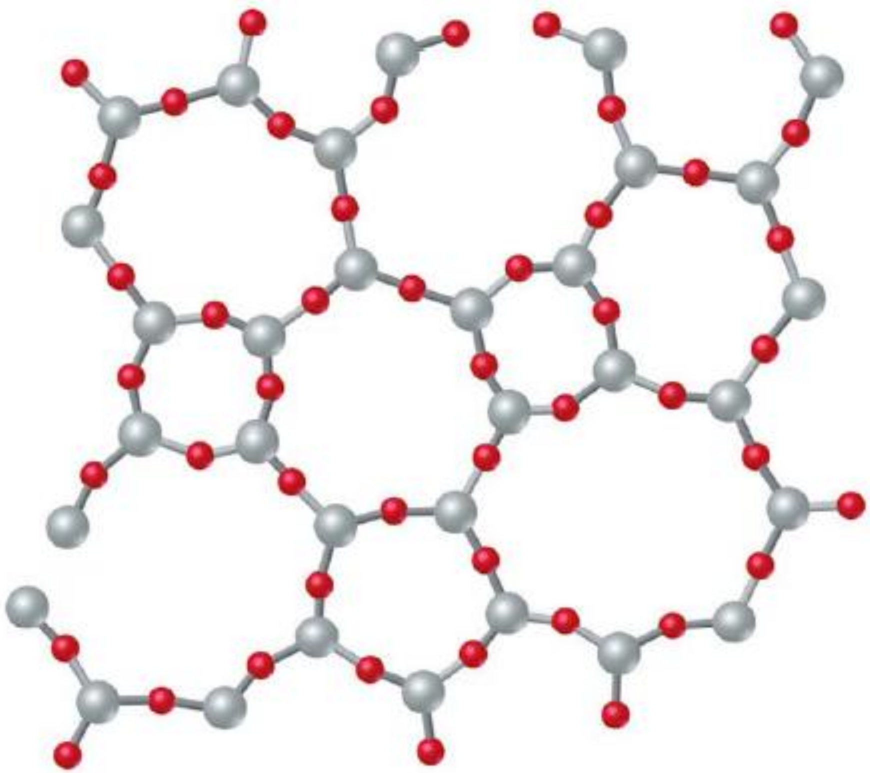
Solids

- We can think of solids as falling into two groups:
  - Crystalline—particles are in highly ordered arrangement.



Although these structures appear to be planar in the drawing, they are actually in a tetrahedral arrangement.

Solids

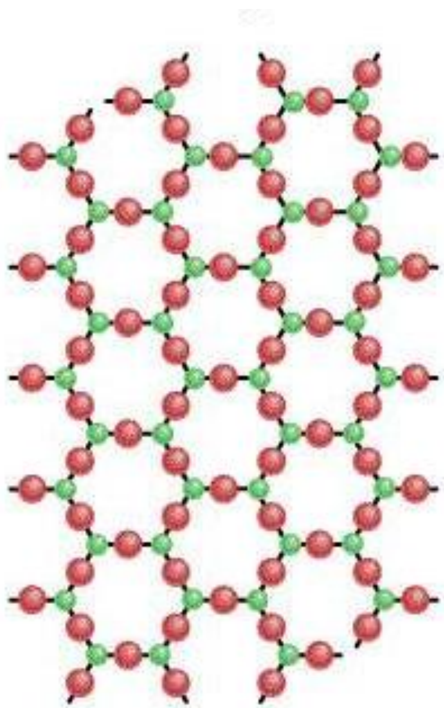


Amorphous SiO_2

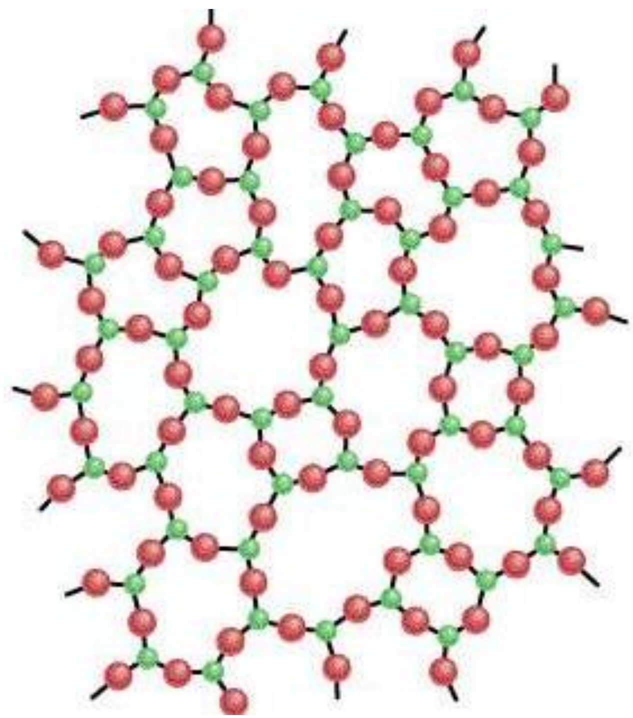
- Amorphous—no particular order in the arrangement of particles.

An ***amorphous solid*** does not possess a well-defined arrangement and long-range molecular order.

A ***glass*** is an optically transparent fusion product of inorganic materials that has cooled to a rigid state **without crystallizing**



Crystalline quartz (SiO_2)

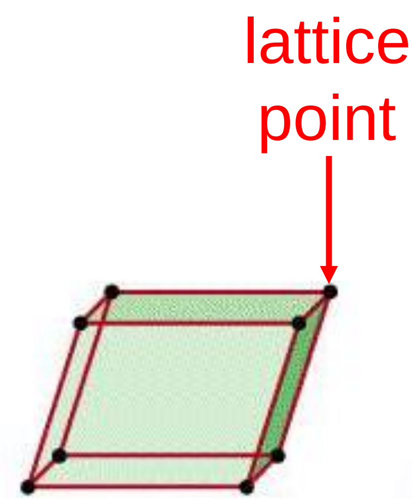


Non-crystalline quartz glass

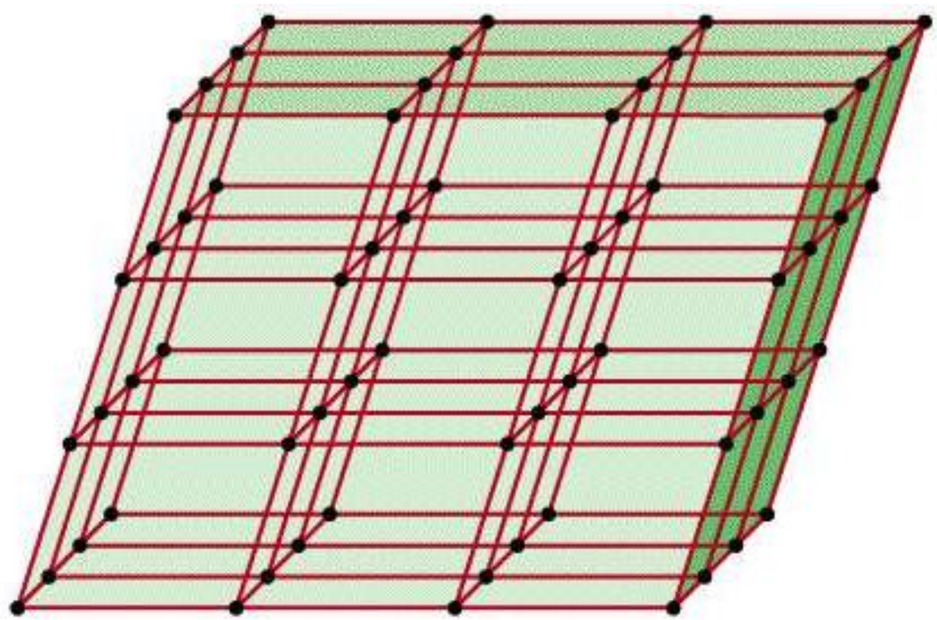
A ***crystalline solid*** possesses rigid and long-range order. In a crystalline solid, atoms, molecules or ions occupy specific (predictable) positions.

An ***amorphous solid*** does not possess a well-defined arrangement and long-range molecular order.

A ***unit cell*** is the basic repeating structural unit of a crystalline solid.



Unit Cell

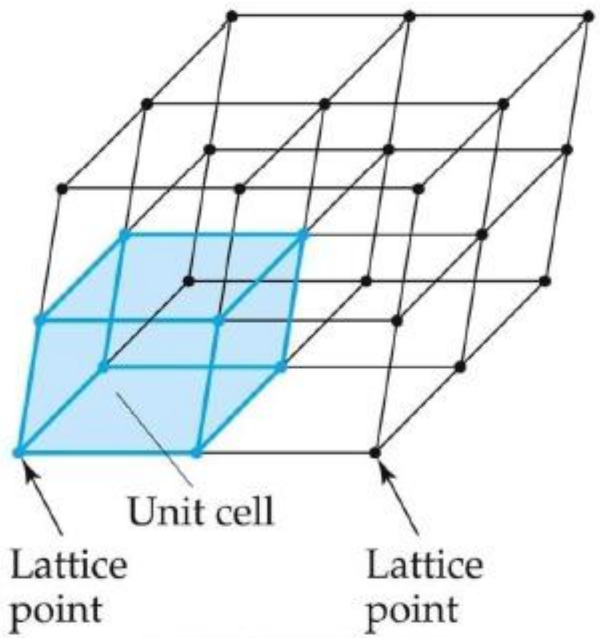


Unit cells in 3 dimensions

<u>At lattice points:</u>

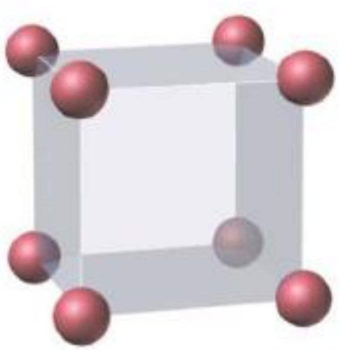
- Atoms
- Molecules
- Ions

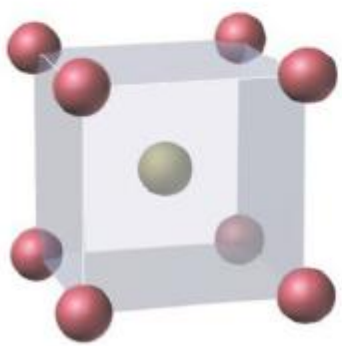
Crystalline Solids

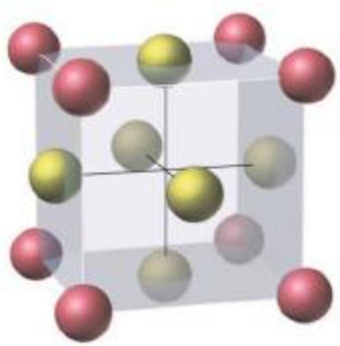


Because of the order in a crystal, we can focus on the repeating pattern of arrangement called the **unit cell**.

Some variations of a cubic unit cell are diagrammed below.


Primitive cubic

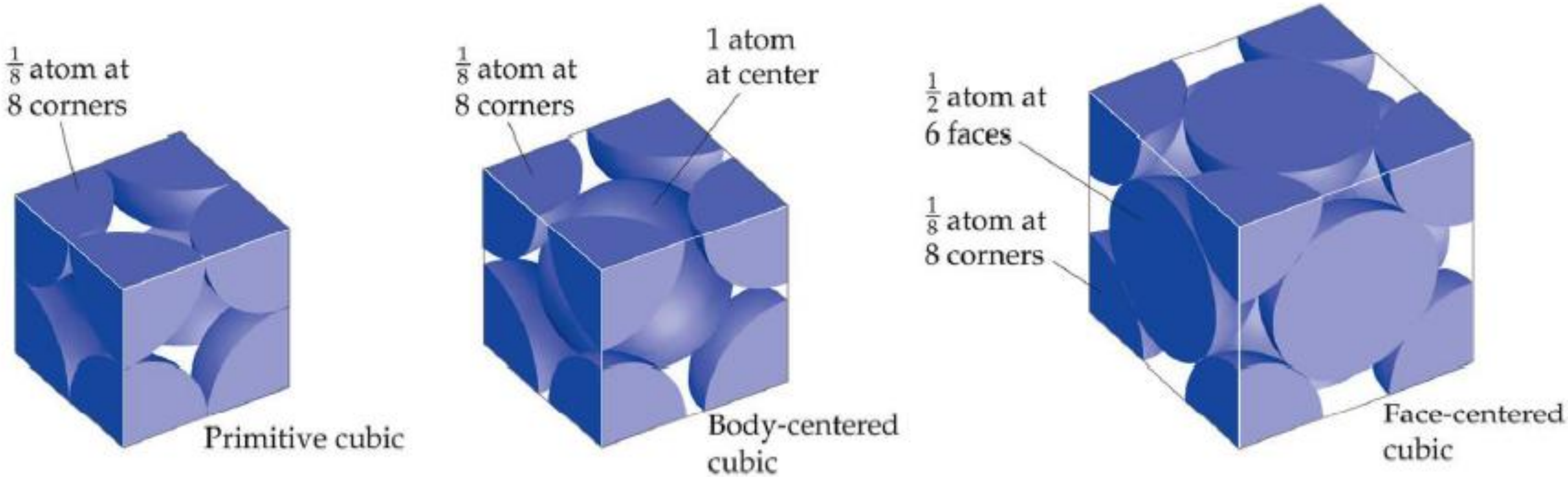

Body-centered cubic


Face-centered cubic

Go to Glencoe animations chapter 11 part I and II packing spheres (units cells)

http://glencoe.mcgraw-hill.com/sites/0023654666/student_view0/chapter11/animations_center.html

The Cubic Unit Cell

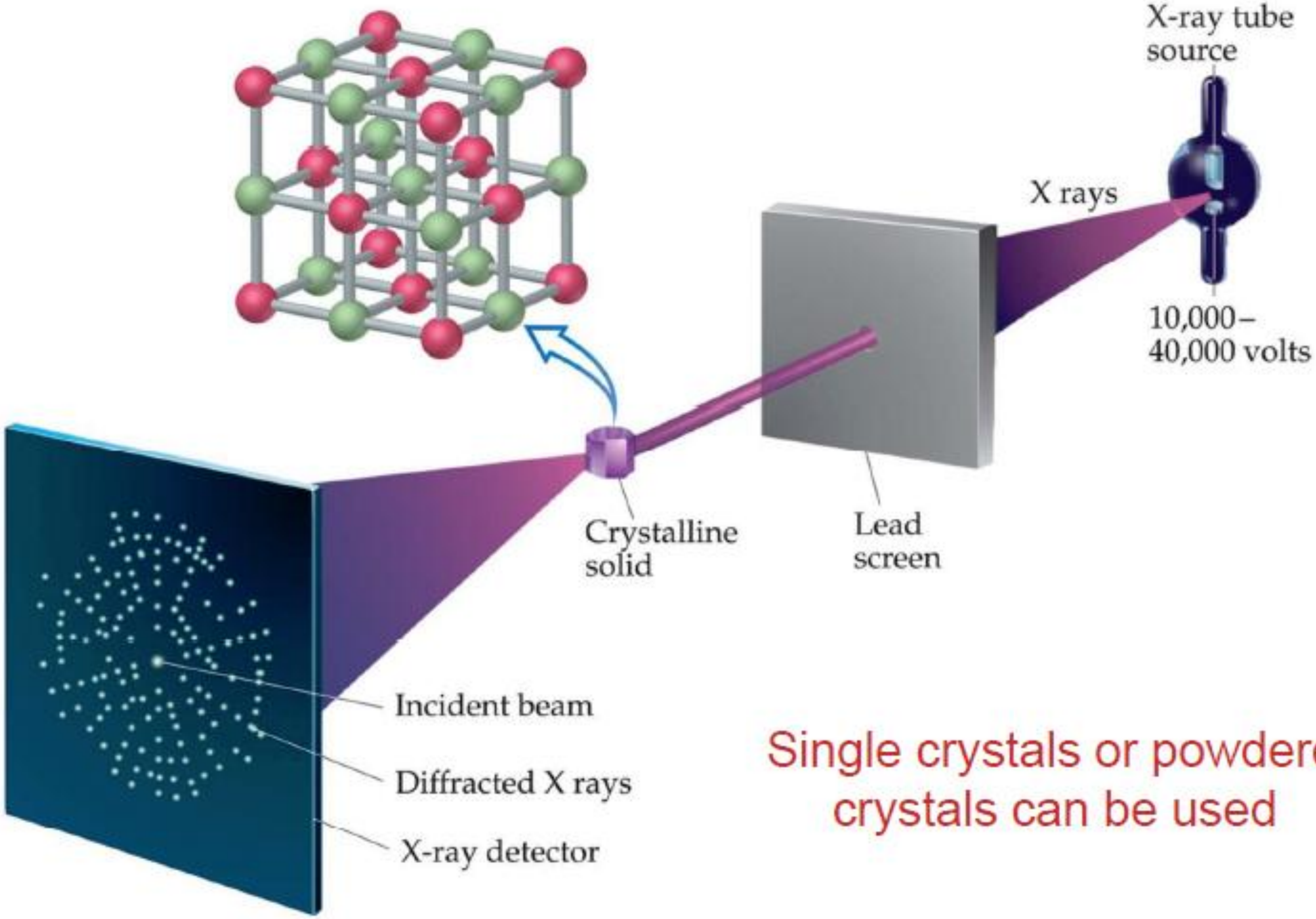


Lattice points in a unit cell are considered to be at the nuclei of the atoms making up the unit cell.

A simple (or primitive) unit cell contains one atom.

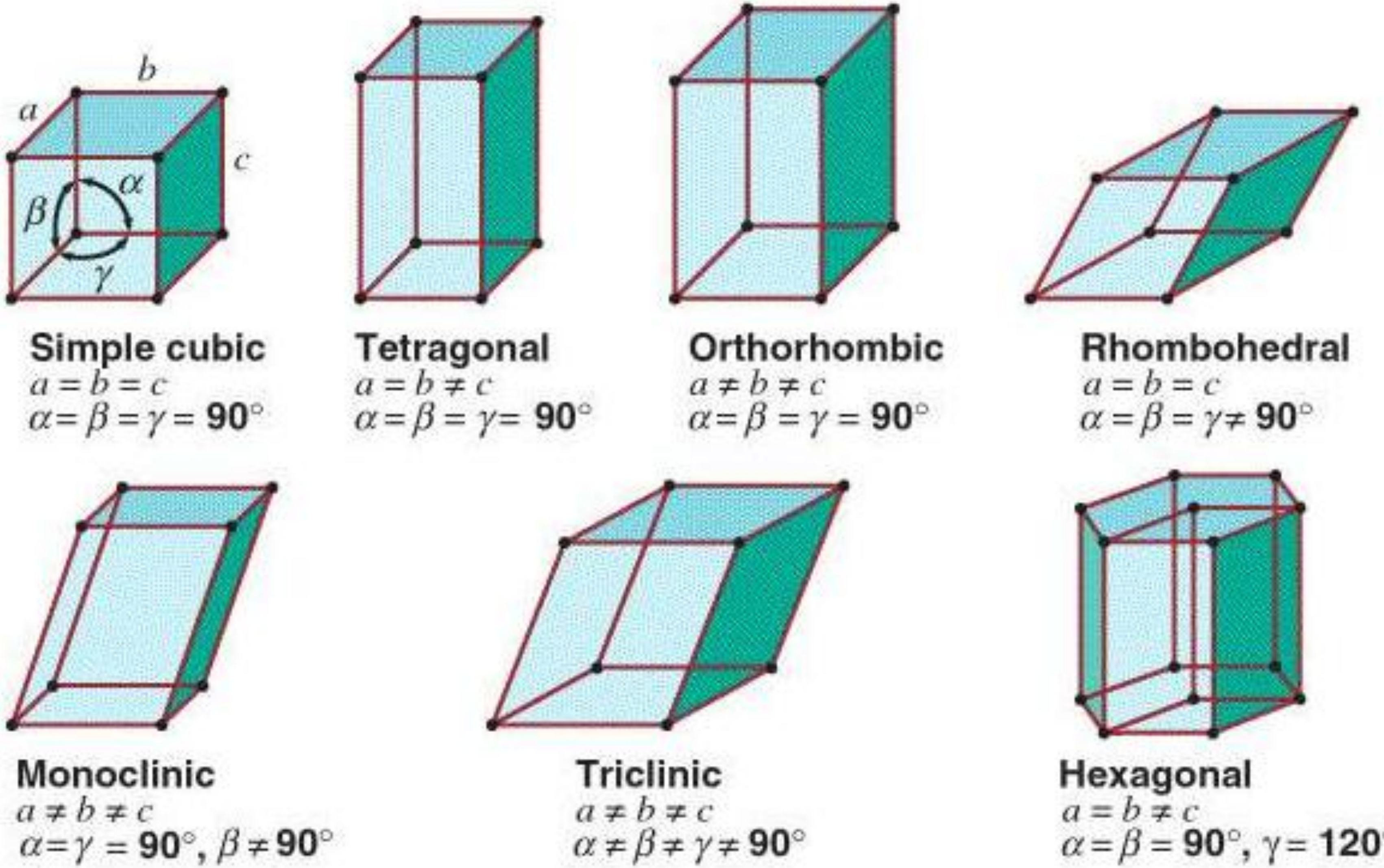
Body-centered and face-centered unit cells contain two or more atoms enclosed in the unit cell.

Unit cell structures are determined by x-ray crystallography



Single crystals or powdered crystals can be used

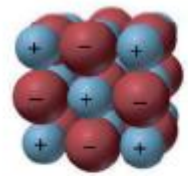
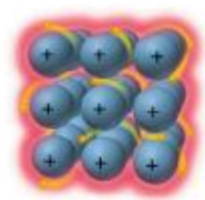
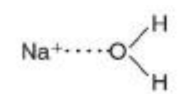
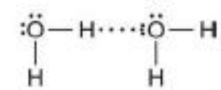
Seven Types of Unit Cells



Types of Bonding in Crystalline Solids

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

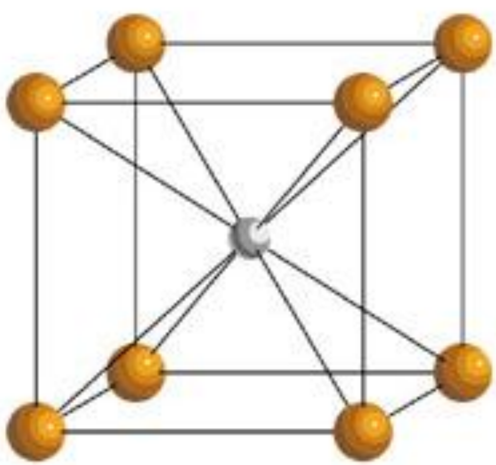
Types of Bonding in Crystalline Solids

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
Nonbonding (Intermolecular)				
Ion-dipole		Ion charge–dipole charge	40–600	
H bond	δ^- δ^+ δ^- —A—H·······:B—	Polar bond to H–dipole charge (high EN of N, O, F)	10–40	
Dipole-dipole		Dipole charges	5–25	I—Cl····I—Cl
Ion–induced dipole		Ion charge–polarizable e^- cloud	3–15	$Fe^{2+}\cdots O_2$
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

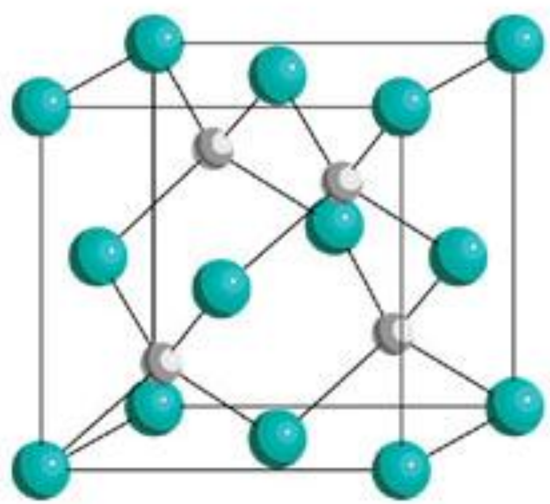
Types of Crystals

<u>Ionic Crystals</u> – Ion-Ion interactions are the strongest (including the "intermolecular forces" (H bonding, etc.)

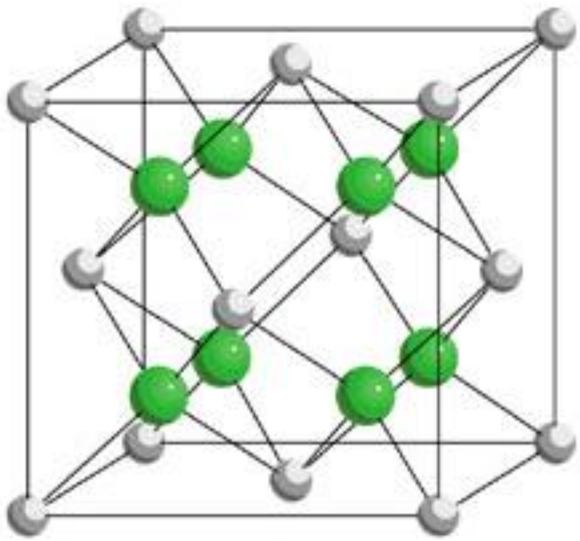
- Lattice points occupied by cations and anions
- Held together by electrostatic attraction
- Hard, brittle, high melting point
- Poor conductor of heat and electricity



CsCl



ZnS



CaF_2

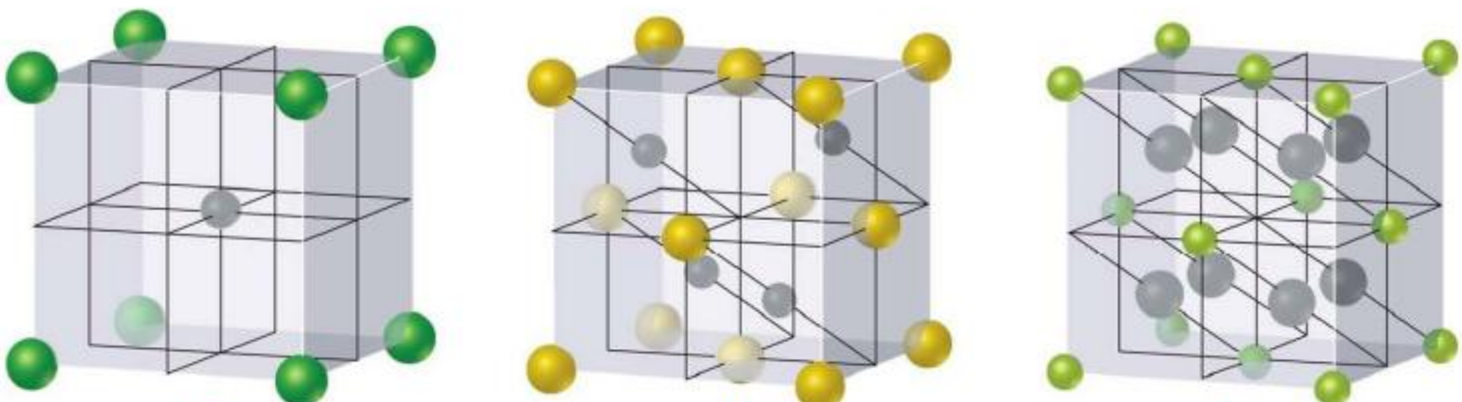
Ionic Solids

What are the empirical formulas for these compounds?

(a) Green: chlorine; Gray: cesium

(b) Yellow: sulfur; Gray: zinc

(c) Green: calcium; Gray: fluorine

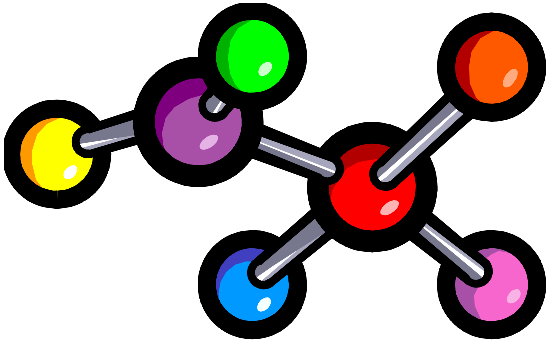


(a) CsCl | (b) ZnS | (c) CaF_2

Types of Crystals

Molecular Crystals

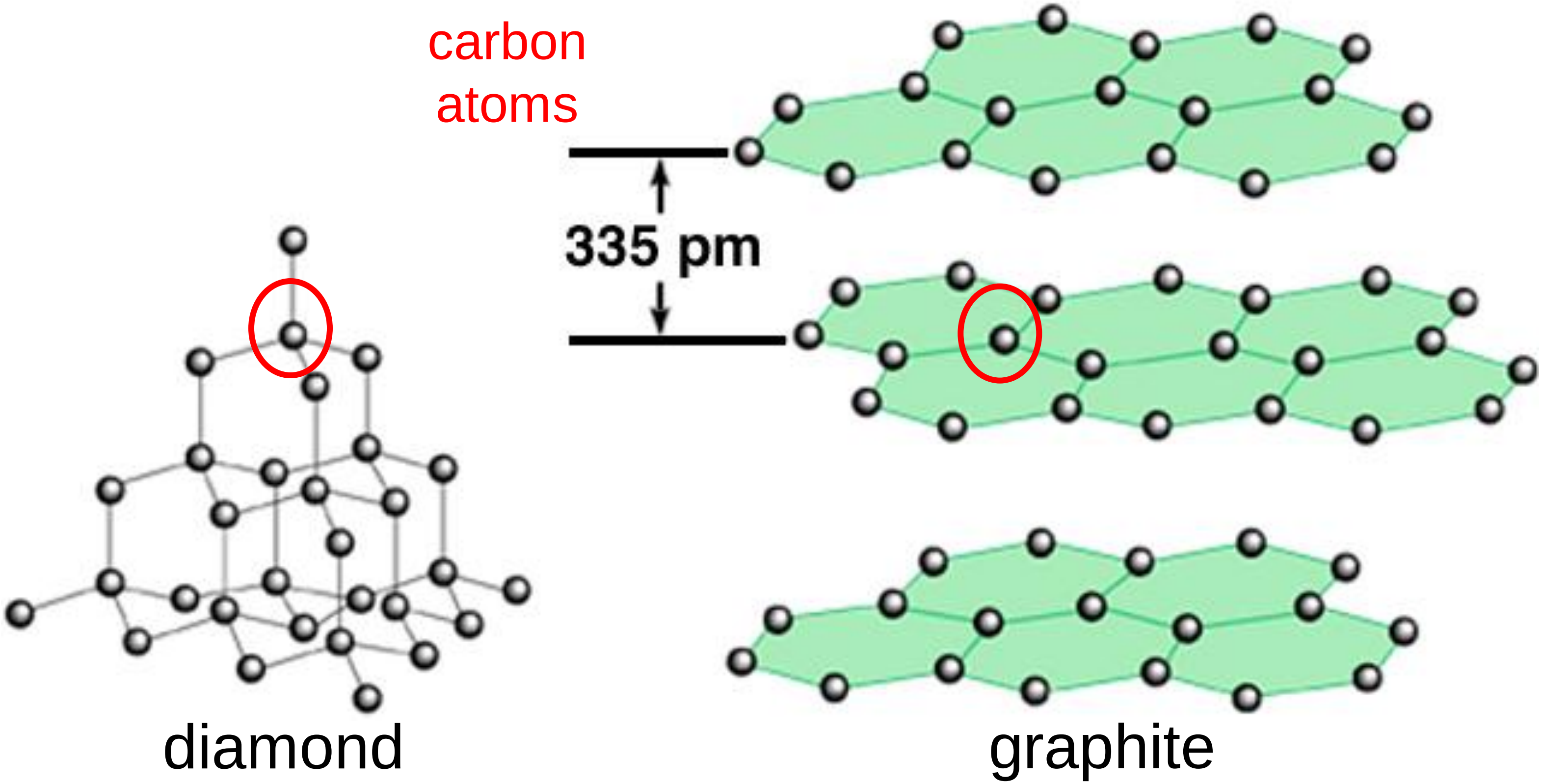
- Lattice points occupied by molecules
- Held together by intermolecular forces
- Soft, low melting point
- Poor conductor of heat and electricity



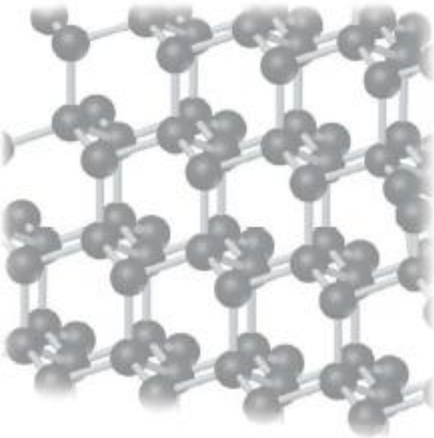
Types of Crystals

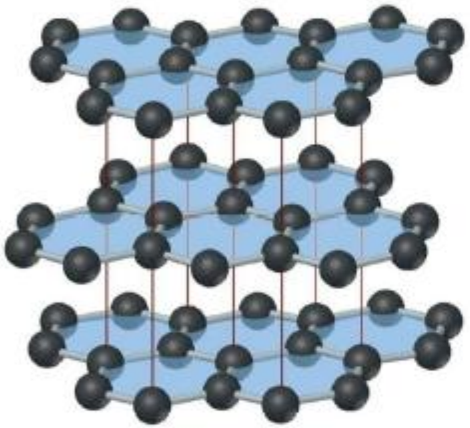
Covalent Crystals – Stronger than IM forces but generally weaker than ion-ion

- Lattice points occupied by atoms
- Held together by covalent bonds
- Hard, high melting point
- Poor conductor of heat and electricity



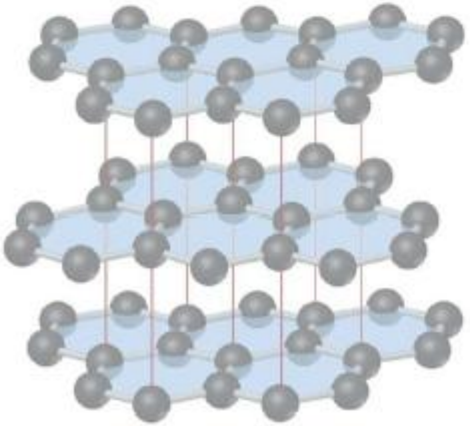
Covalent-Network and Molecular Solids





- Graphite is an example of a molecular solid in which atoms are held together with van der Waals forces.
  - They tend to be softer and have lower melting points.

Covalent-Network and Molecular Solids



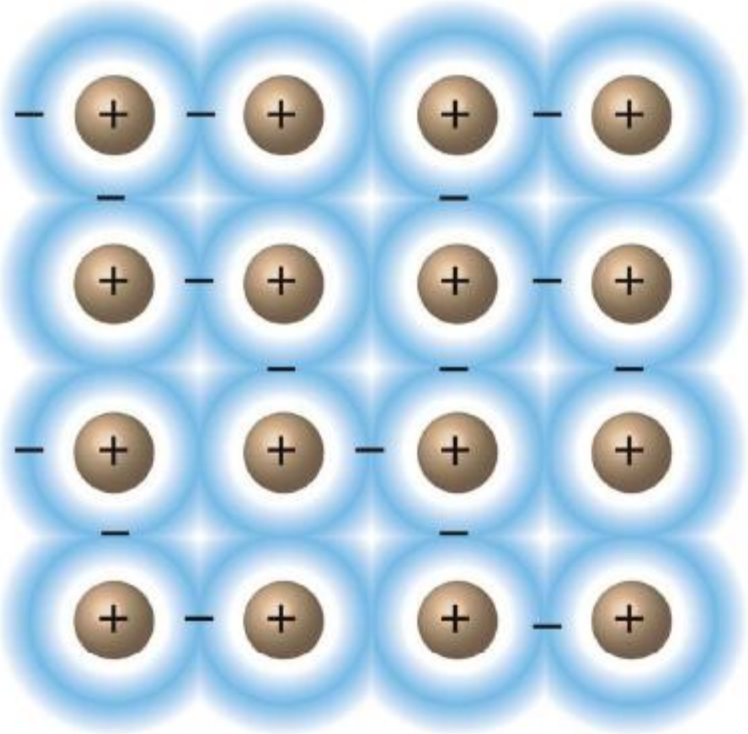
- Diamonds are an example of a covalent-network solid in which atoms are covalently bonded to each other.
  - They tend to be hard and have high melting points.

Types of Bonding in Crystalline Solids

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

Metallic Solids

- Metals are not covalently bonded, but the attractions between atoms are too strong to be van der Waals forces.
- In metals, valence electrons are delocalized throughout the solid.

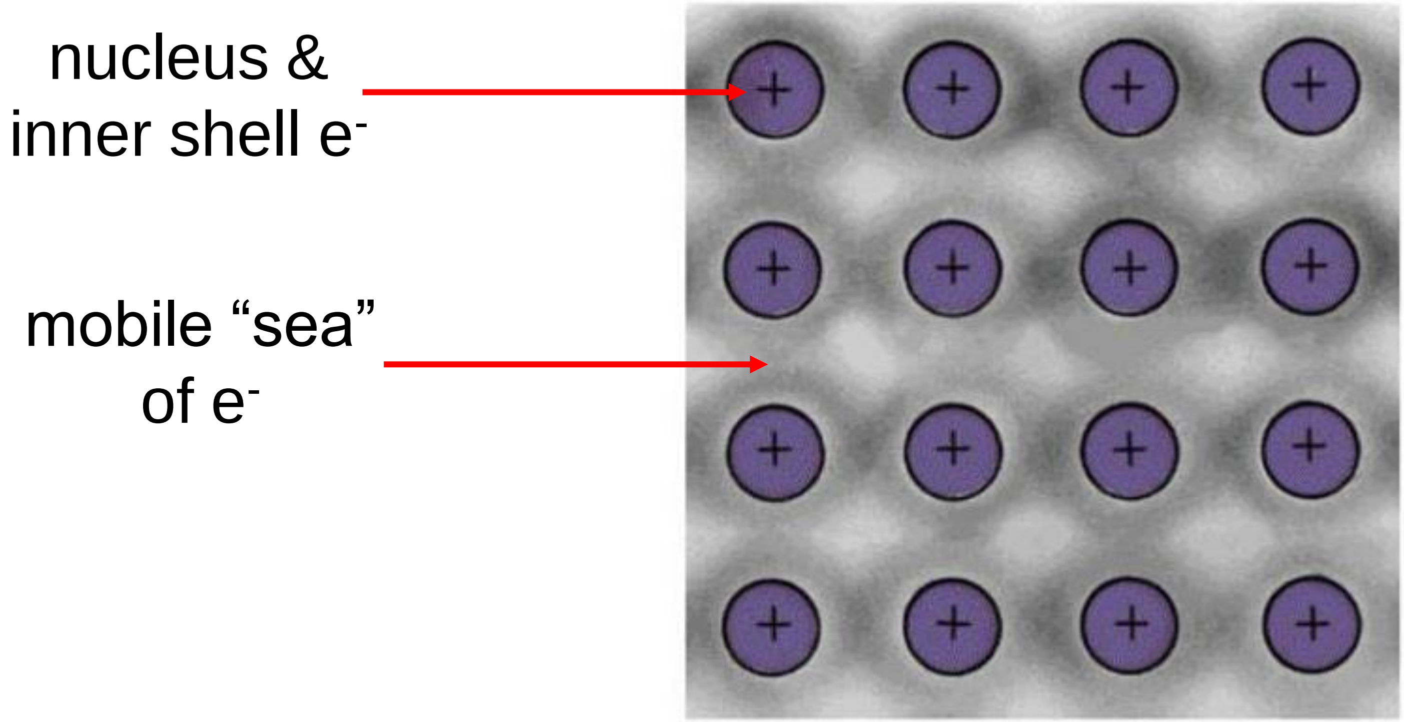


Types of Crystals

Metallic Crystals – Typically weaker than covalent, but can be in the low end of covalent

- Lattice points occupied by metal atoms
- Held together by metallic bonds
- Soft to hard, low to high melting point
- Good conductors of heat and electricity

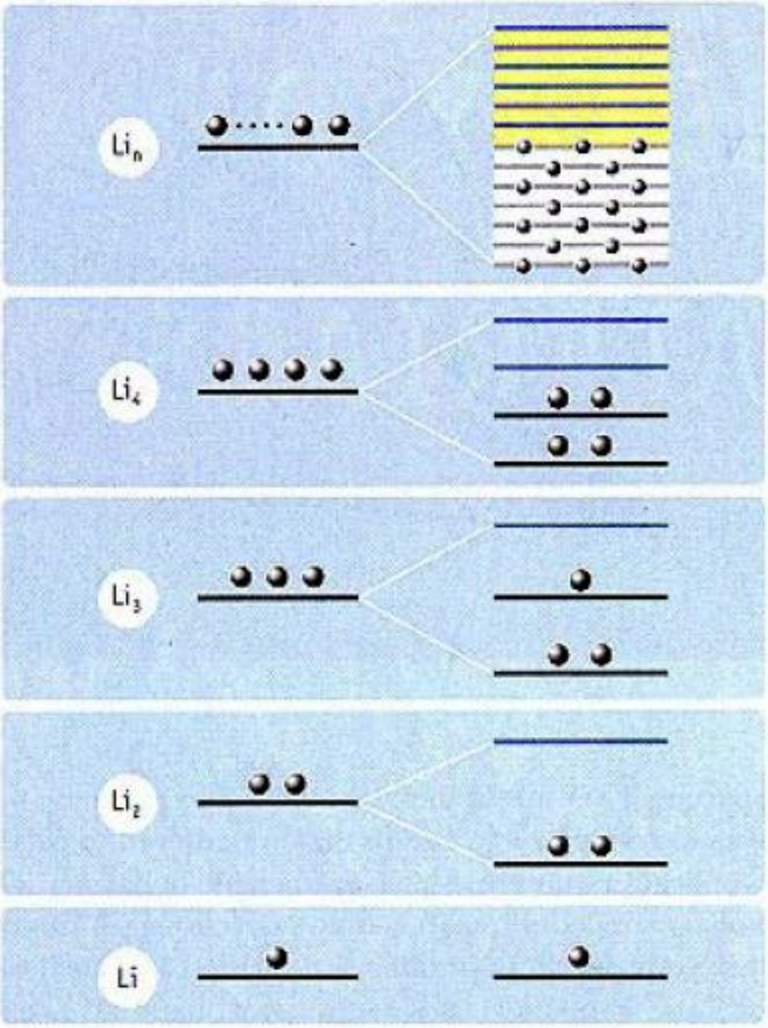
Cross Section of a Metallic Crystal



Molecular Orbitals in Lithium

The 2s orbitals in lithium atoms combine to form molecular orbitals.

In a metallic solid, the orbitals are so close, they merge to form bands of molecular orbitals



Depending on the solid material, the bands can be continuous or separated