Chemistry 3A

Introductory General Chemistry

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- Properties of Liquids and Solids as Phases
- Surface Tension, Viscosity, Capillary Action
- Crystalline Solids
- Energy, Heat, Temperature
- Heat Capacity
- Describing Transitions of Phases: Melting, Freezing, Sublimation, Boiling, Evaporation, Condensation

Liquid, Solids, Gases: Properties

- Liquids and solids: condensed phases because particles in contact

Solids

- particles in fixed positions
- Definite shape and volume
- Usually hard (crystals/rock), but sometimes soft (fat/wax)

Figure 7.1.1: A crystalline arrangement of quartz crystal cluster. Some large crystals look the way they do because of the regular arrangement of atoms (ions) in their crystal structure. (Source: Wikipedia.)

- Ionic solids quite brittle:
  3-D array of positive & negative ions (crystal)
- Large molecule solids (glass) cannot organize particles as crystals → amorphous solids

Liquid, Solids, Gases: Properties

Liquids

- particles have enough (thermal) energy to overcome intermolecular interactions, but particles still move while contacting each other
- Definite volume, but no definite shape

Figure 7.1.2: The formation of a spherical droplet of liquid water minimizes the surface area, which is the natural result of surface tension in liquids. (Source: Wikipedia.)

Liquid, Solids, Gases: Properties

Gases

- Like liquids, particles have enough (thermal) energy to overcome intermolecular interactions & separate from each other, moving randomly in space
- NO definite shape or volume
- Volume increases by a 1000 times or more

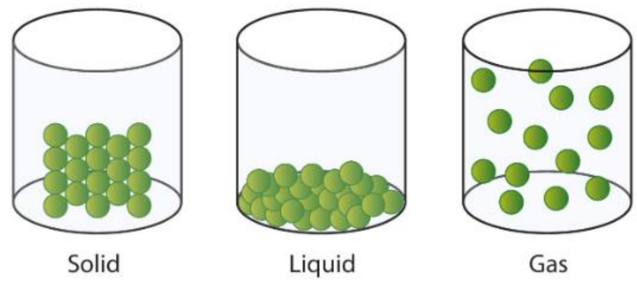


Figure 7.1.3: A Representation of the Solid, Liquid, and Gas States. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

Table 7.1.1: Characteristics of the Three States of Matter

Characteristic	Solid	Liquid	Gas
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

Water

- The liquid of life, the “universal solvent”
- The density of ice is lower than for water
- Water also absorbs and releases energy (as heat) without large changes in temperature, unlike solid metals like steel, made up of element iron (Fe)

Substance	Melting Point	Boiling Point
water	0°C	100°C
ammonia	−78°C	−33°C
Methane	−182°C	−162°C

Some Terms, Definitions

- In the topics on surface tension, viscosity and capillary action to be discussed, certain terms should be understood
- Cohesion describes the force or attraction of molecules of the same kind/identity to each other. This describes how H_2O (water) molecules will create a water droplet. Verb infinitive: to cohere
- Adhesion describes the force or attraction of molecules of different kind/identity to each other, such as when H_2O (water) sticks to sides of glass. Verb infinitive: to adhere

Surface Tension

- Surface tension is the property of a liquid's surface to resist an external force by minimizing its surface area, caused by the cohesive (intermolecular) forces between liquid molecules that are imbalanced at the surface
- Water has a high surface tension: it explains why water forms droplets form on a waxy surface (leaves, car bodies) rather than spread themselves thin
- Minimizing surface area: explains why water forms spherical droplets because sphere is smallest possible surface area for any volume

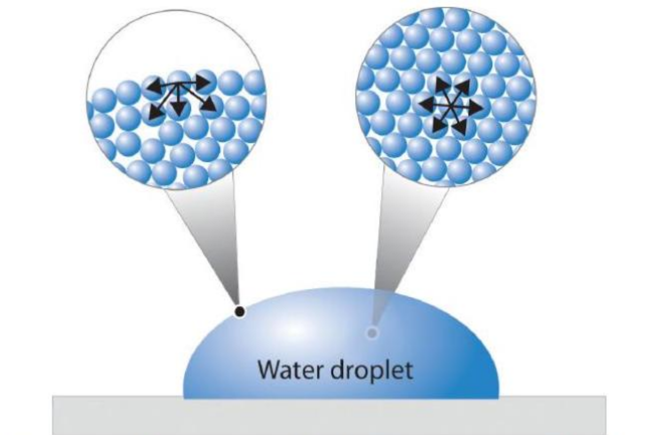


Figure 7.2.1: A Representation of Surface Tension in a Liquid. Molecules at the surface of water experience a net attraction to other molecules in the liquid, which holds the surface of the bulk sample together. In contrast, those in the interior experience uniform attractive forces.

Surface Tension UNITS

- Surface tension is a quantity with units of joules (J) per square meter (m^2): J / m^2 OR
  dyne (dyn) per centimeter (cm): dyn/cm
- The joule is measure of **energy** while the dyne and the newton [N] (1 N = 100,000 dyn) is a measure of **force**
- The higher (stronger) the intermolecular forces, the higher the surface tension
- Water has very high intermolecular force (because of hydrogen bonding [later] while organic molecules have lower intermolecular force/surface tension

Force (F) is related to energy (E) as work (W) (both work and energy are measured in joules) by a distance/length (d) factor

$$E \text{ or } W = F \times d$$

You learn this is physics

Surface Tension Fun

- Surface tension at zero gravity: check it out
- https://www.youtube.com/watch?v=lMtXfwk7PXg

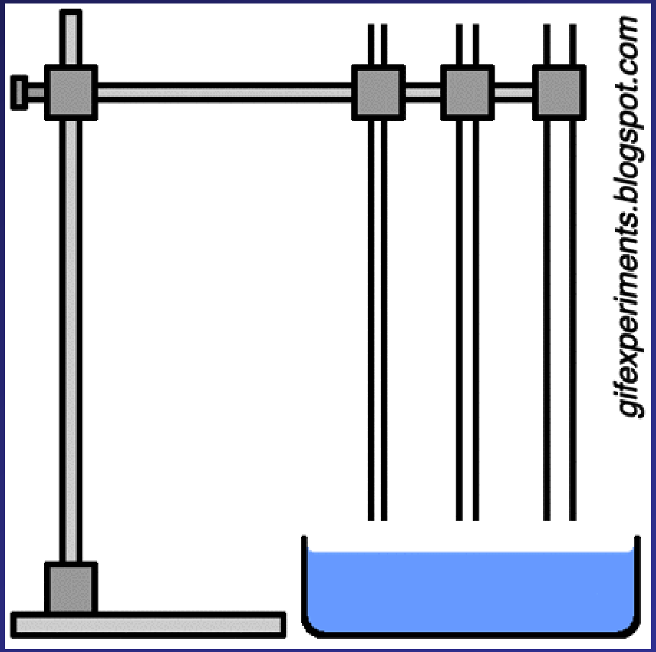
Capillary Action

- Capillary action is the movement of a liquid through or along a solid material, even against gravity, driven by the forces of cohesion (a liquid's attraction to itself), adhesion (a liquid's attraction to a solid surface), and surface tension.

- When the adhesive forces between the liquid and the solid are stronger than the cohesive forces within the liquid, the liquid will "climb" the solid, which is seen in examples like a paper towel absorbing a spill or water traveling up a plant's xylem

Capillary Action

- It is called “capillary action” because a capillary tube, which is a very small diameter tube, when one of its open ends is placed in a liquid (water, even blood), will draw a small sample volume of that liquid into the tube as a result of adhesion to the glass and cohesion with other liquid molecules



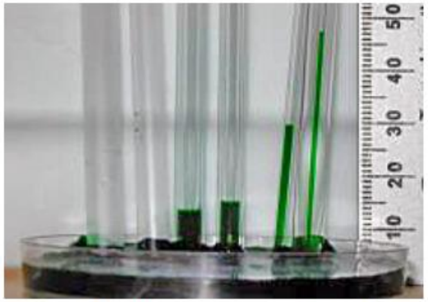
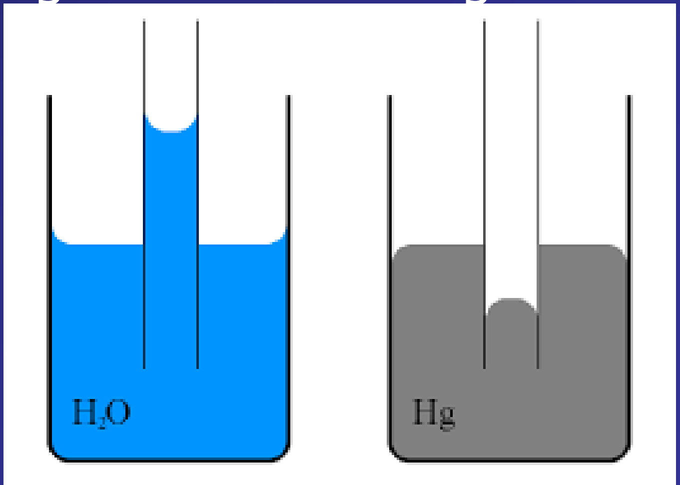


Figure 7.2.3: The Phenomenon of Capillary Action. Capillary action seen as water climbs to different levels in glass tubes of different diameters. Credit: Dr. Clay Robinson, PhD, West Texas A&M University.

Water vs Mercury in Capillary Action

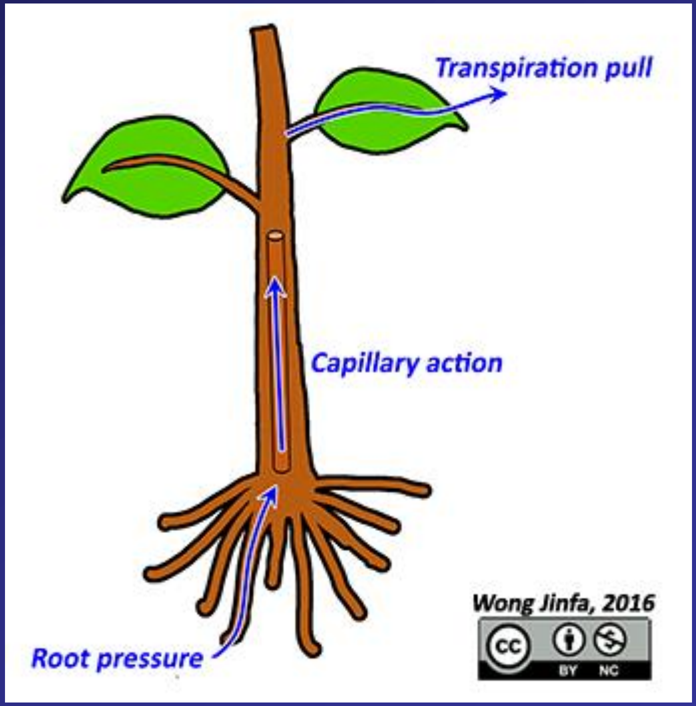
- Water will have strong adhesion to glass. Glass is composed of silanol (–Si–OH) groups (note the H-bonding hydroxyls), which interact with water's natural H-bonding hydroxyls (H–O–H), creating a force which draws a volume of water above the level of water in a container although. The meniscus in the capillary is concave

- Mercury however has no strong bonding interaction to glass wall silanol groups, so adhesion is nil. However, the cohesive forces in mercury are much stronger and they pull any mercury in the capillary down below the level of the mercury in the container. The meniscus in the capillary is convex



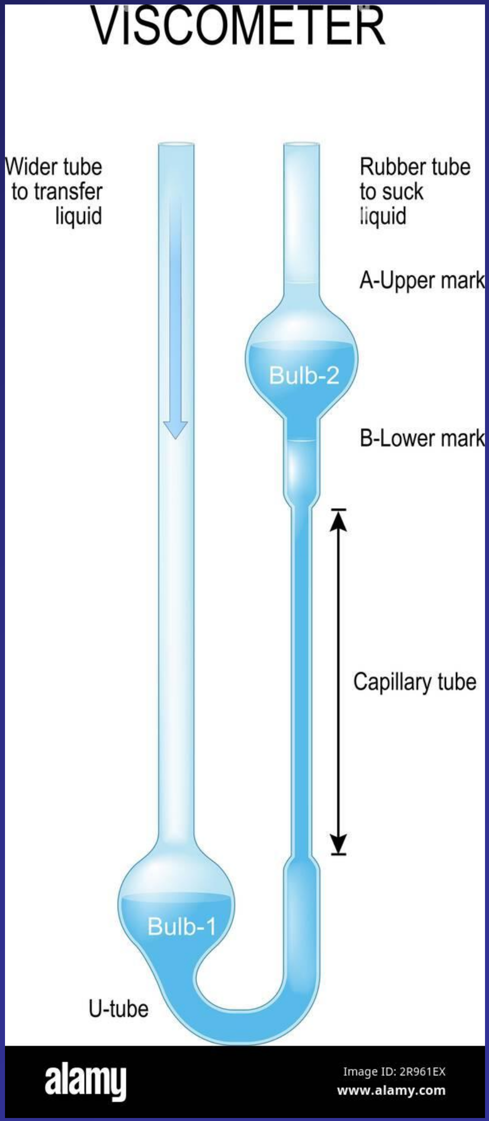
Biological Capillary Action

- Capillary action is also observed in plant life
- Fluids and nutrients are transported up stems and tree trunks as water adheres easily in the capillary-like structure of the plant’s xylem



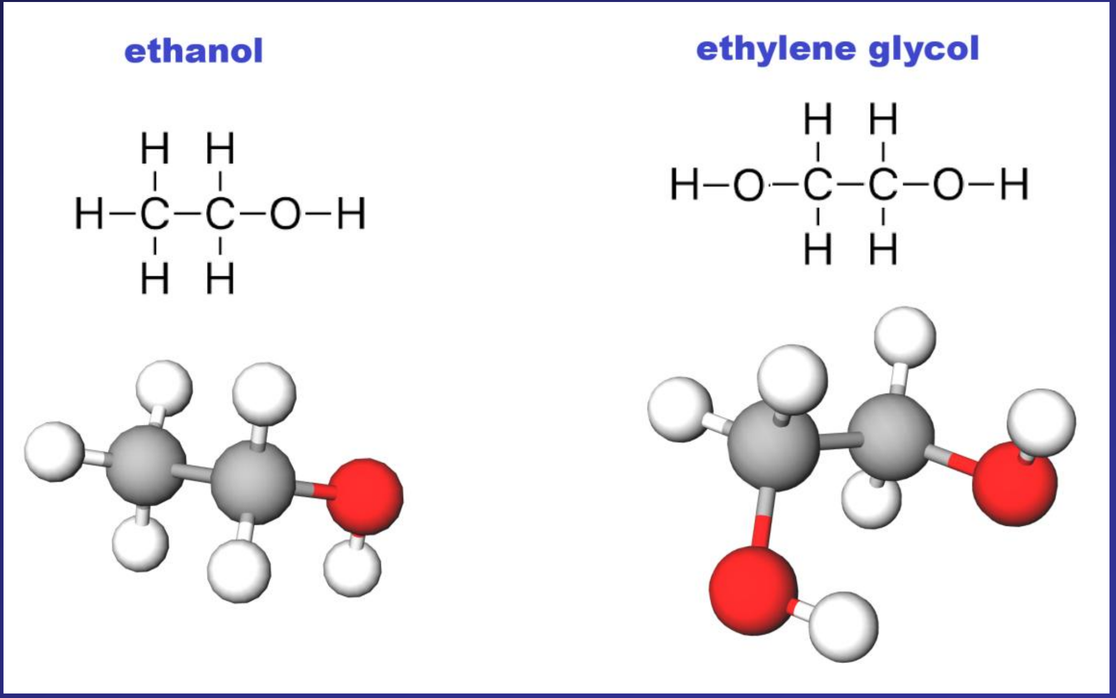
Viscosity

- Viscosity (symbol Greek letter *eta* η) is the resistance of a fluid to flow, often described as its "thickness" or internal friction
- A viscometer is a special glass device that measures time it takes for test liquid to flow through narrow vertical tubing
- Units of viscosity: poise (1 poise = 1 mPa s)
- If a liquid has strong intermolecular forces, the molecules will not move past each other but slowly. Adding an –OH group to ethanol (CH_3CH_2OH) to make ethylene glycol ($HOCH_2CH_2OH$)



Viscosity

- If a liquid has strong intermolecular forces, the molecules will not move past each other but slowly. Adding an –OH group to ethanol (CH_3CH_2OH) to make ethylene glycol ($HOCH_2CH_2OH$) increases viscosity by 15 times, as it creates a second interconnecting bridge-like point through hydrogen bonding



Liquid Properties in Review

- The table reveals patterns (correlations) in surface tension and stronger intermolecular forces
- Although mercury (Hg) does not show the strong intermolecular forces usually in hydrogen bonding, it does have a high surface tension indicative of metallic bonding

Table 7.2.1: Surface Tension, Viscosity, Vapor Pressure (at 25°C Unless Otherwise Indicated), and Normal Boiling Points of Common Liquids

Substance	Surface Tension ($\times 10^{-3}$ J/m^2)	Viscosity (mPa•s)	Vapor Pressure (mmHg)	Normal Boiling Point (°C)
		Organic Compounds		
diethyl ether	17	0.22	531	34.6
n-hexane	18	0.30	149	68.7
acetone	23	0.31	227	56.5
ethanol	22	1.07	59	78.3
ethylene glycol	48	16.1	~0.08	198.9
		Liquid Elements		
bromine	41	0.94	218	58.8
mercury	486	1.53	0.0020	357
		Water		
0°C	75.6	1.79	4.6	—
20°C	72.8	1.00	17.5	—
60°C	66.2	0.47	149	—
100°C	58.9	0.28	760	—

Motor Oil

Crystalline Solid Classes

There are four classes of crystalline solids

1. Ionic

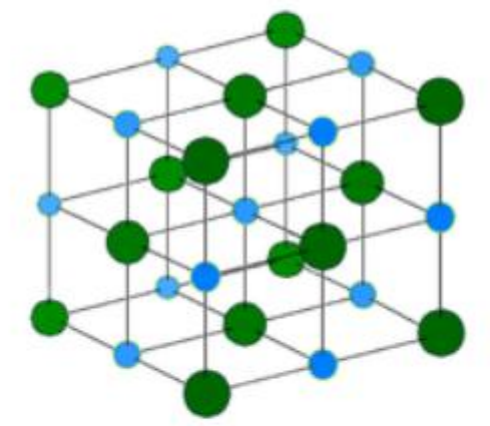


Figure 7.3.1: NaCl crystal.

- Sodium chloride ($NaCl$) is the classic monatomic type, but polyatomic types also exist. The formula unit in the crystal shows alternating cations & anion
- Typically Group 1 & 2 cations combine with Group 16 & 17 anions
- Crystals hard, brittle, high melting points
- They are not electrically conductive as solids, but in aqueous solution and molten (liquid) state, they are conductive

Crystalline Solid Classes

2. Metallic

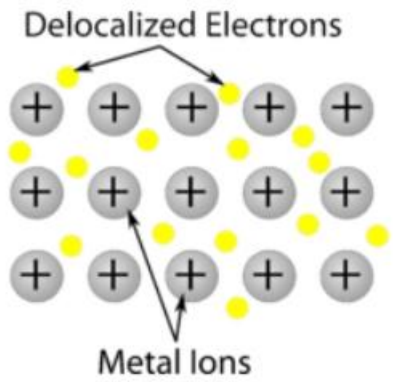


Figure 7.3.2: Metallic crystal lattice with free electrons able to move among positive metal atoms.

- A metallic crystal is actually metal atoms as cations (positively charged atoms) whose electrons exist as a "lake" or "sea" of valence electrons not tethered or bound to the metal atoms (delocalized electrons). This makes these solid metals excellent electrical conductors as electrons move freely
- Melting points are HIGHER for metals when:
  - More valence electrons -- Al (3) > Mg (2) > Na (1)
  - Higher ionic charge pulls on delocalized electrons – Al^{3+} > Mg^{2+} > Na^{+} (direct consequence of previous point)
  - Smaller ionic radius – stronger electrostatic attraction between cation and the delocalized mobile electrons

Crystalline Solid Classes

3. Covalent Network

- All atoms in the crystal form a very strong covalent bond to neighboring atoms usually in an orderly lattice
- Classic example is diamond, composed of carbons bonded in tetrahedral shape in a large network
- Other examples are pure silicon (Si), quartz (SiO_2), carborundum (SiC), borazon (BN)
- The atoms are not ions (ionic), so they are NOT electrically conductive

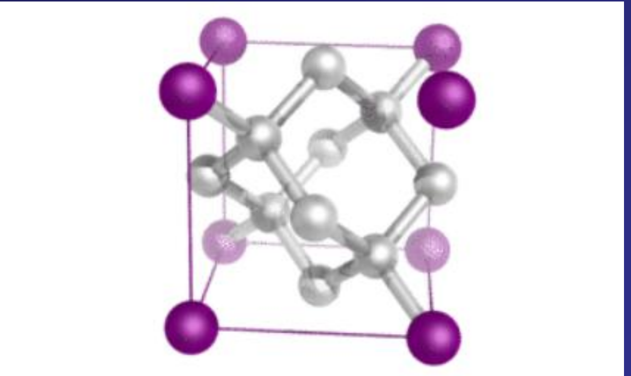
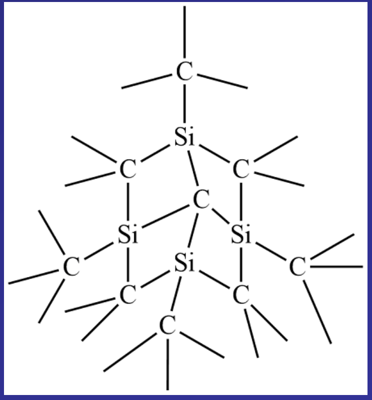
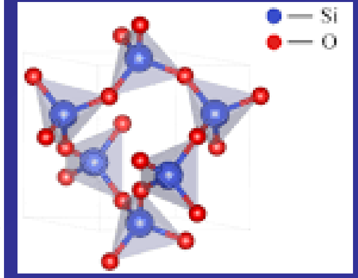


Figure 7.3.3: Diamond is a network solid and consists of carbon atoms covalently bonded to one another in a repeating three-dimensional pattern. Each carbon atom makes for single covalent bonds in a tetrahedral geometry.





Crystalline Solid Classes

4. Molecular

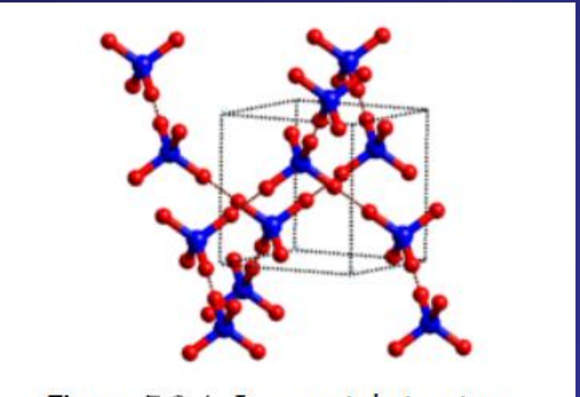


Figure 7.3.4: Ice crystal structure.

- Previous solids had ions, metal cations, and atoms in networks
- These solids are molecules connected (bonded) by INTERMOLECULAR forces: hydrogen bonding, dipole-dipole, dispersion forces in nonpolar crystals
- Examples: H_2O solid (ice), iodine (I_2), CO_2 solid (dry ice)
- Physical properties like melting & boiling points are lower. Since they are not ionic or having mobile electrons, they are poor electrical conductors

Review: Crystalline Solid Classes

Problem solving:

Classify Ge, RbI, $C_6(CH_3)_6$, Zn, CO_2, $BaBr_2$, GaAs, AgZn as ionic, molecular, covalent network, metallic and order them by melting points

(exercises in your book)

Table 7.3.2: Properties of the Major Classes of Solids

Ionic Solids	Molecular Solids	Covalent Solids	Metallic Solids
poor conductors of heat and electricity	poor conductors of heat and electricity	poor conductors of heat and electricity*	good conductors of heat and electricity
relatively high melting point	low melting point	high melting point	melting points depend strongly on electron configuration
hard but brittle; shatter under stress	soft	very hard and brittle	easily deformed under stress; ductile and malleable
relatively dense	low density	low density	usually high density
dull surface	dull surface	dull surface	lustrous

*Many exceptions exist. For example, graphite has a relatively high electrical conductivity within the carbon planes, and diamond has the highest thermal conductivity of any known substance.

Table 7.3.1: Crystalline Solids: Melting and Boiling Points

Type of Crystalline Solid	Examples (formulas)	Melting Point (°C)	Normal Boiling Point (°C)
Ionic	$NaCl$	801	1413
	CaF_2	1418	1533
Metallic	Hg	-39	630
	Na	371	883
	Au	1064	2856
	W	3410	5660
Covalent Network	B	2076	3927
	C (diamond)	3500	3930
	SiO_2	1600	2230
Molecular	H_2	-259	-253
	I_2	114	184
	NH_3	-78	-33
	H_2O	0	100

Universe: System & Surroundings

- Forms of Energy: kinetic, potential
  Transferred as heat or work
  Always conserved (not created or destroyed) in the universe
- The universe
  - The system: what we are looking at (observing) or studying in experiment
  - The surroundings: everything outside or not part of the system
  - Energy exchange/transfer occurs between system and surroundings: energy lost by system is gained by surroundings and vice-versa

Endothermic / Exothermic Reactions

- Endothermic reactions are when heat/energy comes or is absorbed from surrounding into system

  (endo- "within", "inside", "taking in") (-thermic is heat or energy)

- Exothermic reactions are when heat/energy is released from system into surroundings

  (exo- "out", "outward") (-thermic is heat or energy)

- When phase changes occur, energy changes also happen
- Dry ice vaporization means CO_2 molecules absorbs energy
- Water becoming ice means H_2O molecules release energy to surroundings

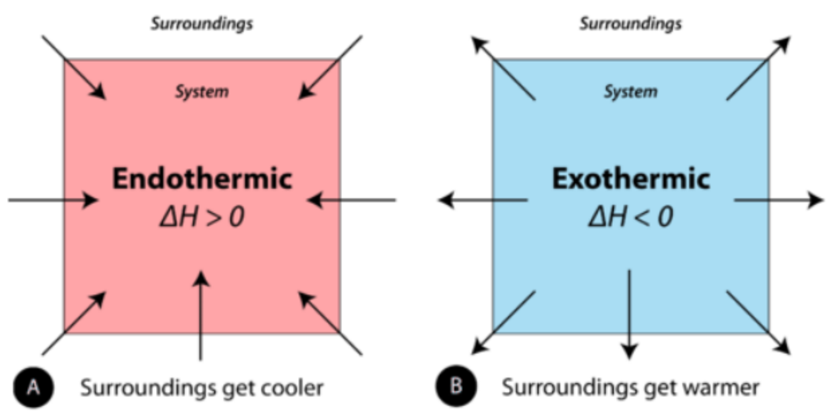


Figure 7.4.1.1: (A) Endothermic reaction. (B) Exothermic reaction.

Endothermic reaction: surroundings get cooler and delta H is greater than 0, Exothermic reaction: surroundings get warmer and delta H is less than 0

Endothermic or Endothermic?

- Water boiling?
  endothermic
- Gasoline burning?
  exothermic
- Ice forming on pond?
  exothermic
- Water vapor condensing?
  exothermic
- Gold melting?
  endothermic
- The point is to understand if energy is going into the system (from the surroundings) or being removed/released from the system (into the surroundings)

Book Confusion (not for exam)

- In talking about endo- and exothermic processes in reaction with energy transfer between surroundings and system, your book uses the symbol ***q*** and ***ΔH***.
- The proper symbol for explaining energy in this context is ***q***, which is usually energy as heat
- The symbol ***H*** is for enthalpy, which is actually the sum of the internal energy (E) (as heat) of a system AND the work energy related to changes in pressure and volume (PV). Thus $H = E + PV$, and $\Delta H = \Delta E + \Delta(PV)$. When pressure is constant, then $\Delta H = \Delta E$ because $\Delta(PV) = 0$. This point is beyond the scope of this course, but it is essential to be detailed here.

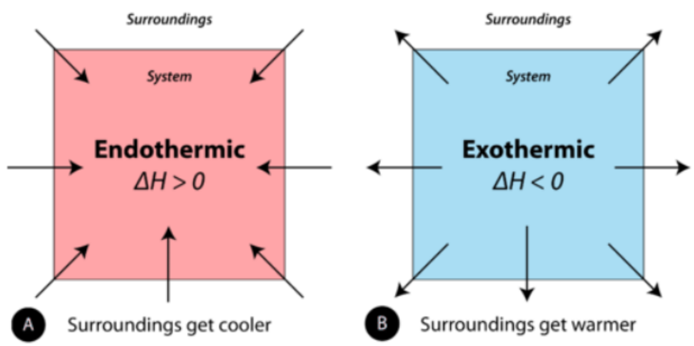


Figure 7.4.1.1: (A) Endothermic reaction. (B) Exothermic reaction.

Endothermic reaction: surroundings get cooler and delta H is greater than 0, Exothermic reaction: surroundings get warmer and delta H is less than 0

Energy: Heat & Work

- Energy in chemistry/physics comes in the form of **heat** (symbolized ***q***)
- It can be observed using temperature measurements
- Energy is also in form of **work.** The classic physics equation of work is of a **force** applied over a **distance**

$$\boldsymbol{W = F \times d}$$

Energy: Kinetic and Potential

- **Kinetic energy** is the energy of **motion**. It's the energy that an object possesses due to its movement. The more massive an object is and the faster it moves, the more kinetic energy it has

Figure 7.4.3.2: A wind farm in Solano County harnesses the kinetic energy of the wind. (CC BY-SA 3.0 Unported; BDS2006 at Wikipedia)

- **Potential energy** is **stored energy** 🔋. It's the energy an object has due to its **position** or **composition**, waiting to be released or converted into another form of energy, like kinetic energy

Energy: Chemical Potential

- **Chemical Potential Energy** is the stored potential energy in atoms, molecules, and the chemical bonds between them
- This includes the positions between particles of matter (atoms, molecules)
- It also includes the composition of the substance (compound) which affects why atoms and molecules are positioned with respect to each other

Measuring Energy

- Energy is a quantity, so it has a **number** with **units**
- Units are the **calorie** (**cal**) and the **joule** (**J**)
- Definition of calorie: the amount of energy (heat) needed to raise 1 gram H_2O (water) by 1°C
- Capital "C" calorie: the calories in nutrition food are a different measure:

  1 Calorie = 1 kilocalorie = 1000 calories
- To get joules, use **4.184 J = 1 cal** as conversion factor!

Heat Capacity

- **Heat capacity** (***C***) is the amount of **heat energy** (***q***) required to raise temperature of a substance by one degree Celsius (°C) or one Kelvin (K)
- **Specific Heat capacity** is the amount of required to raise 1 gram of a substance by 1°C

For water, it is 1 cal/g×°C (or 4.184 J/g×°C)

- To calculate energy (heat, ***q***) transferred to/from a mass, the specific heat, the temperature change, and the mass are used in the calculation

$$q = c_p \times m \times \Delta T$$

If the final T is greater than initial T, both *q* and ΔT are greater than zero (> 0). If final T is less than initial T, then *q* and ΔT are less than zero (< 0)

Why does swimming pool heat faster than wading pool?

Specific Heat

- The specific heat is a property of a substance or compound. NOTE THE UNITS: energy per amount of mass per temperature
- You will be provided with these numbers and are not expected to memorize them
- Be able to use your knowledge of algebra to make the calculation!

Table 7.5.1.1: *Specific Heats of Some Common Substances*

Substance	**Specific Heat** $(J/g^{o}C)$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

Calculations

A 15.0 g piece of cadmium (Cd) metal absorbs of 134 J heat while rising from 24.0°C to 62.7°C. Calculate the specific heat of cadmium

Needed: $q = c_p \times m \times \Delta T$

Solving for c_p? $c_p = \frac{q}{m \times \Delta T}$ this is algebra

$$c_p = \frac{134\ \mathrm{J}}{15.0\ \mathrm{g} \times (62.7°\mathrm{C}\ - 24.0°\mathrm{C})} = 0.231 \frac{\mathrm{J}}{\mathrm{g}\ °\mathrm{C}}$$

Note we had 3 significant digits

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Needed: $q = c_p \times m \times \Delta T$

Solving for q? Need c_p for iron from table: $c_p = 0.108 \frac{\mathrm{cal}}{\mathrm{g}\ °\mathrm{C}}$

$$q = 150.0\ \mathrm{g} \times \frac{0.108\ \mathrm{cal}}{\mathrm{g}\ °\mathrm{C}} \times (73.3°\mathrm{C}\ - 25.0°\mathrm{C}) = 782\ \mathrm{cal}$$

The heat flows into the metal since temperature is increasing

Thoughts

Explain what happens when heat flows into or out of a substance at its melting point or boiling point

- The energy goes into changing the phase, not the temperature

How does the amount of heat required for a phase change relate to the mass of the substance?

- The amount of heat is a constant per gram of substance

Phase Transitions

- The table shows the phase transitions of matter and the (name of) the process associated with the transition
- Any temperature, like melting and boiling points, known for a substance/compound with a phase transition is for the pure substance
- These processes are **isothermal**: that means the temperature does not change while energy (heat) is being added or released from the substance!

Process	**Phase Transition**
melting	solid → liquid
freezing/solidification	liquid → solid
boiling/vaporization/evaporation	liquid → gas
condensation	gas → liquid
sublimation	solid → gas
deposition/desublimation	gas → solid

Curiosities of Phases

- A liquid is observed as a fills a container. It exists at its temperature with the substance/compound molecules or atoms brought together by intermolecular forces
- The liquid fills a container because of gravity (another type of force)

- But in a zero-gravity environment, it still retains its phase as a liquid because of those intermolecular forces based on other physical properties

(a) (b)

Figure 7.6.1: Liquids and Gravity. (a) A liquid fills the bottom of its container as it is drawn downward by gravity and the particles slide over each other. (b) A liquid floats in a zero-gravity environment. The particles still slide over each other because they are in the liquid phase, but now there is no gravity to pull them down. Source: Photo on the left © Thinkstock. Photo on the right courtesy of NASA, http://www.nasa.gov/mission_pages/st...image_009.html.

Phase Transitions

- Melting, freezing, boiling or evaporation, and other phase transitions occur on a visible-to-the-eye (macroscopic) level
- But at the sub-microscopic level, it is important to note the interatomic/molecular forces that maintain a solid, liquid, or gaseous state

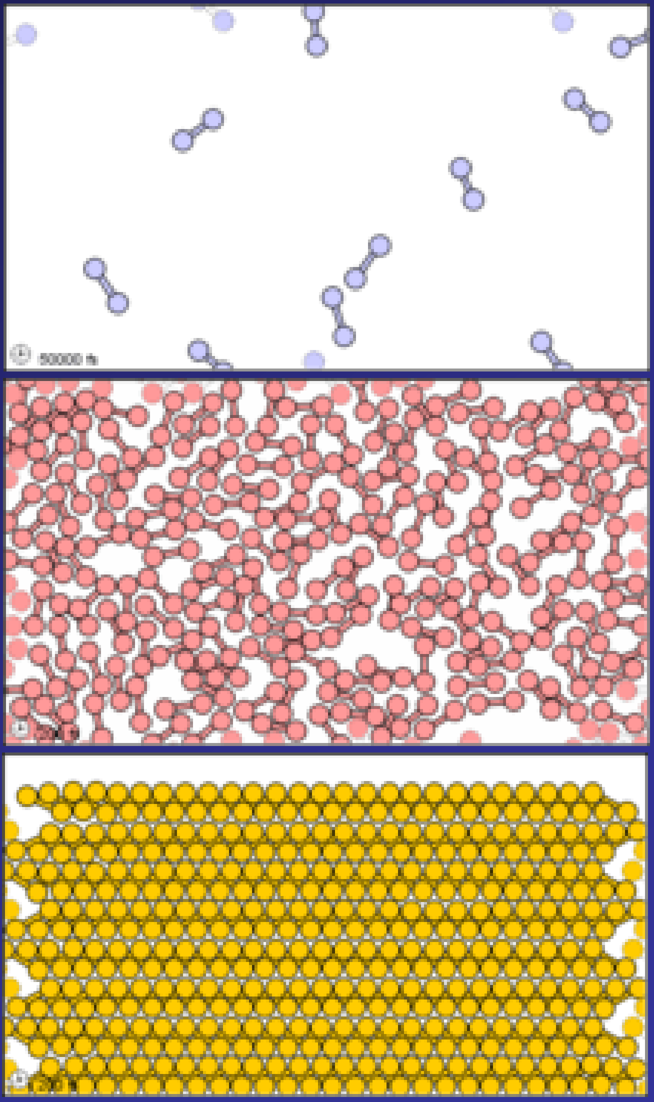


Figure 7.6.2: Sub-microscopic view of the diatomic molecules of the element bromine (a) in the gaseous state (above 58°C); (b) in liquid form (between -7.2 and 58.8°C); and (c) in solid form (below -7.2°C). As a solid, the molecules are fixed, but fluctuate. As a liquid, the molecules are in contact but are also able to move around each other. As a gas, most of the volume is actually empty space. The particles are not to scale; in reality, the dots representing the particles would be about 1/100th of the size depicted.

Melting/Freezing/Sublimation

- Melting is an isothermal process: it occurs without any change in temperature while energy is added to cause the phase change
- Freezing is opposite of melting as to phase, but it too is isothermal

Thermodynamic Symbol	**Meaning**
ΔH_{vap} = heat of vaporization	Energy to cause liquid → gas
ΔH_{fus} = heat of fusion	Energy to cause solid → liquid
ΔH_{sub} = heat of sublimation	Energy to cause solid → gas
Note that $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$	

Table 7.6.1.2: Melting Points of Common Materials

Materials	Melting Point (°C)
Hydrogen	-259
Oxygen	-219
Diethyl ether	-116
Ethanol	-114
Water	0
Pure silver	961
Pure gold	1063
Iron	1538

Figure 7.6.1.3: Freezer burn on a piece of beef. (Public Domain; RolloM.)

Boiling/Evaporation/Condensation

Terminology

- Boiling: liquid to gas at boiling point (isothermal)
- Evaporation: liquid→gas BELOW boiling point
- Vaporization: liquid form of substance becomes gaseous: combines boiling and evaporation terms
- Condensation: gas→liquid process and is opposite of vaporization

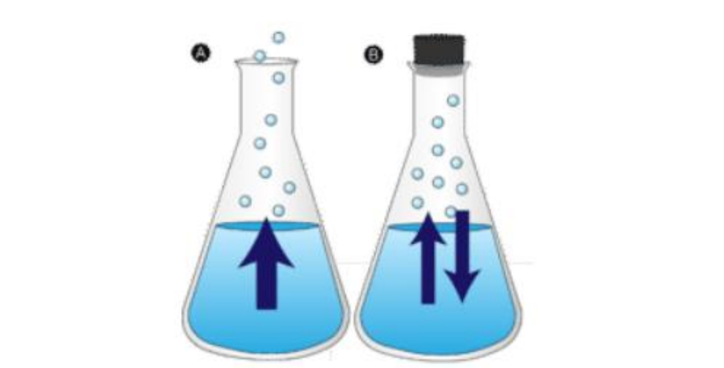


Figure 7.6.2.2: Evaporation (A) and condensation (B).

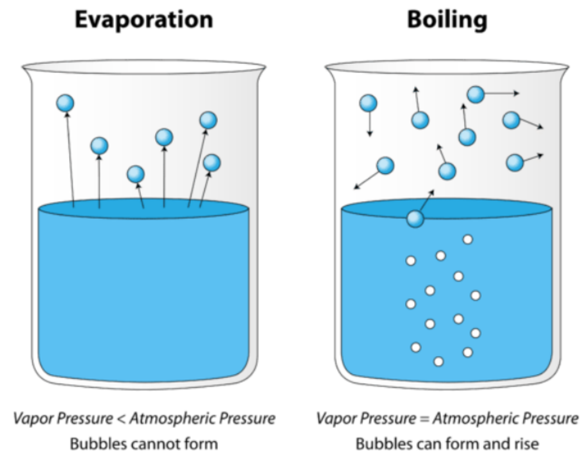


Figure 7.6.2.4: Comparison between evaporation and boiling.

Evaporation Statistics

- A liquid at two temperatures T_1 and T_2, both below boiling points of the liquid, but $T_2 > T_1$
- Plot of # molecules versus energy is done for T_1 and T_2
- The higher temperature T_2 sees faster evaporation because there is an energy point E at which an individual molecule can break free to go from liquid to gas phase, even though not at boiling point

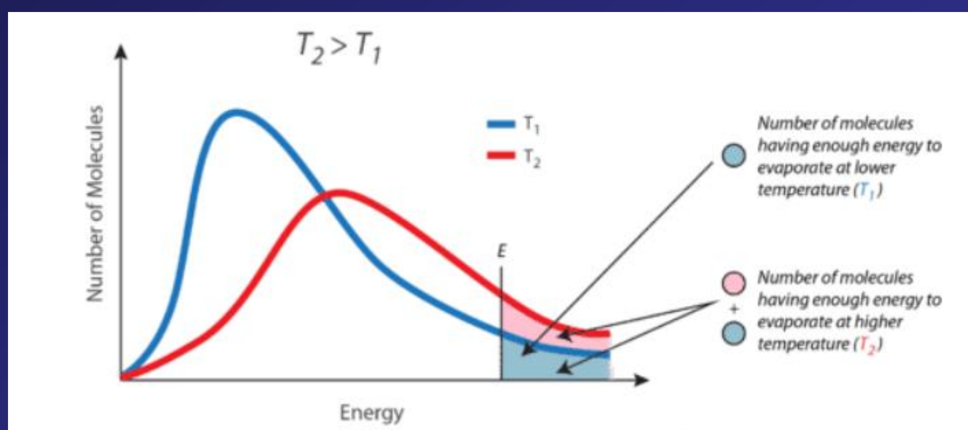


Figure 7.6.2.3: Kinetic energy distribution curves for a liquid at two temperatures T_1 and T_2. The shaded area represents the molecules with enough kinetic energy to escape the liquid and become vapor.

Phase Change Energy Calculations

The total energy (as heat) of a phase transition will be the product of the mass (in moles) and the enthalpy property

$$heat = n \times \Delta H_{fus} \text{ during melting or solidification}$$

$$heat = n \times \Delta H_{vap} \text{ during boiling or condensation}$$

How much energy is needed for 45.7 g H_2O to melt at 0°C?

Find: ΔH_{fus} of water = 6.01 kJ/mol, molar mass H_2O = 18.0 g/mol

Solve: $q = n \times \Delta H_{fus} = \left(45.7 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}}\right) \times \frac{6.01 \text{ kJ}}{1 \text{ mol H}_2\text{O}} = 15.3 \text{ kJ}$

How much energy is needed for 108 g C_6H_6 to FREEZE at 5.5°C?

Find: ΔH_{fus} of benzene=9.9 kJ/mol, molar mass C_6H_6 = 78.11 g/mol

Solve: $q = n \times -\Delta H_{fus} = \left(108 \text{ g C}_6\text{H}_6 \times \frac{1 \text{ mol C}_6\text{H}_6}{78.11 \text{ g C}_6\text{H}_6}\right) \times \frac{-9.9 \text{ kJ}}{1 \text{ mol C}_6\text{H}_6} = -13.7 \text{ kJ}$

Heating Curves

- A **heating curve** (H_2O is shown) is a plot of Temperature (°C) versus Time
- Time is really a measure of Energy Input At A Constant Rate (joules per unit time)
- The curve shows how the solid, liquid and gas phases absorb energy as measured by temperature changes AND it shows energy absorbed during phase changes at constant temperature

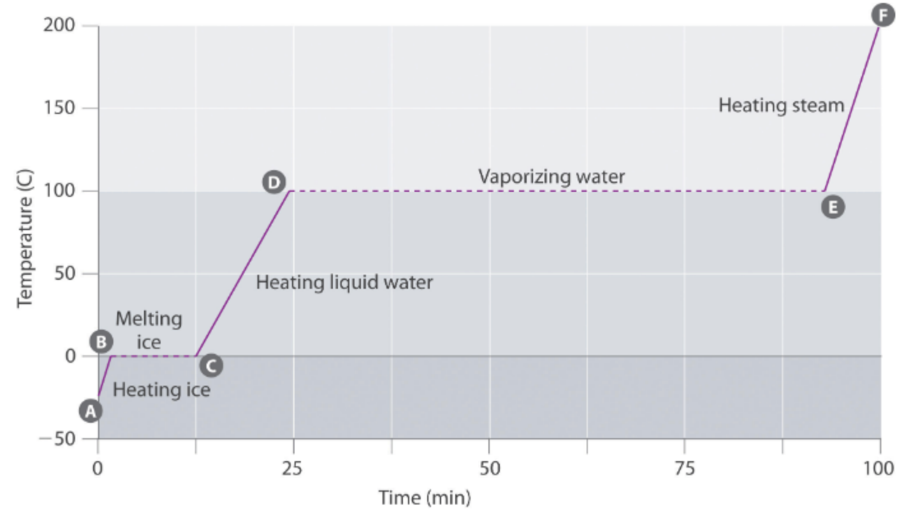


Figure 7.6.4.1: A Heating Curve for Water. This plot of temperature shows what happens to a 75 g sample of ice initially at 1 atm and −23°C as heat is added at a constant rate: A–B: heating solid ice; B–C: melting ice; C–D: heating liquid water; D–E: vaporizing water; E–F: heating steam.

Cooling Curves

- A **cooling curve** (H_2O is shown) is ALSO a plot of Temperature (°C) versus Time
- Time is really a measure of Energy Release At A Constant Rate (joules per unit time)
- The curve shows how the gas (steam), liquid and gas phases release energy as measured by temperature changes AND
it shows energy released during phase changes at constant temperature

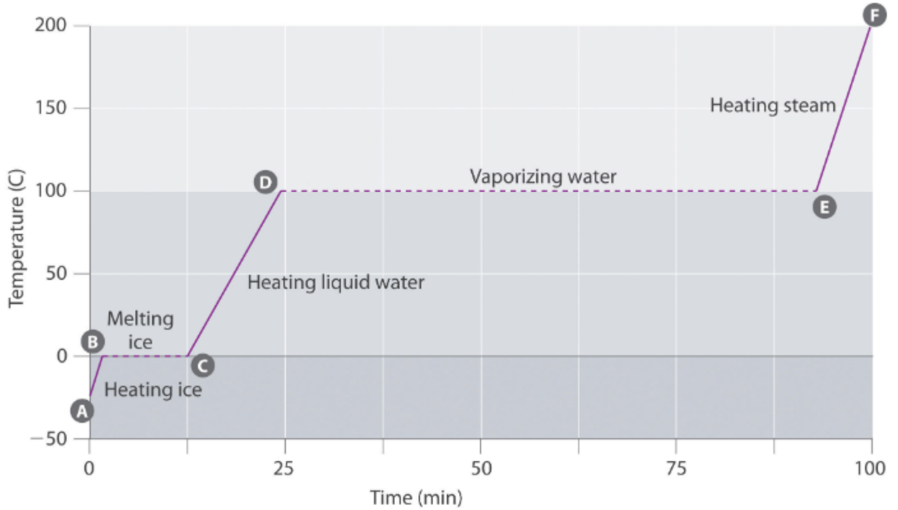


Figure 7.6.4.1: A Heating Curve for Water. This plot of temperature shows what happens to a 75 g sample of ice initially at 1 atm and −23°C as heat is added at a constant rate: A–B: heating solid ice; B–C: melting ice; C–D: heating liquid water; D–E: vaporizing water; E–F: heating steam.

Slide Placeholder for Super-

- This slide to be updated to explain superheating and supercooling