Liquids and Solids

Chapter 13

Water

- Colorless, odorless and tasteless
- Density of ice **<** than density of liquid water
  - Not the normal trend
  - For equal masses of ice & water, ice has the larger volume
  - For equal volumes of ice & water, water has the larger mass
  - Water expands when it freezes
  - Ice floats on water
  - Density of liquid water = 1.00 g/mL
  - Density of ice = 0.917 g/mL

Water

- Freezes at 0°C
  - At 1 atm, solid at 0°C or below
  - Normal freezing point = normal melting point
- Boils at 100°C
  - At 1 atm, liquid up to 100°C, then turns to steam
  - Normal boiling point
  - Boiling point increases as atmospheric pressure increases
- Temperature stays constant during a state change
- Relatively large amounts of energy needed to melt solid or boil liquid
- Liquid’s Specific Heat Capacity = 4.18 J/g-°C

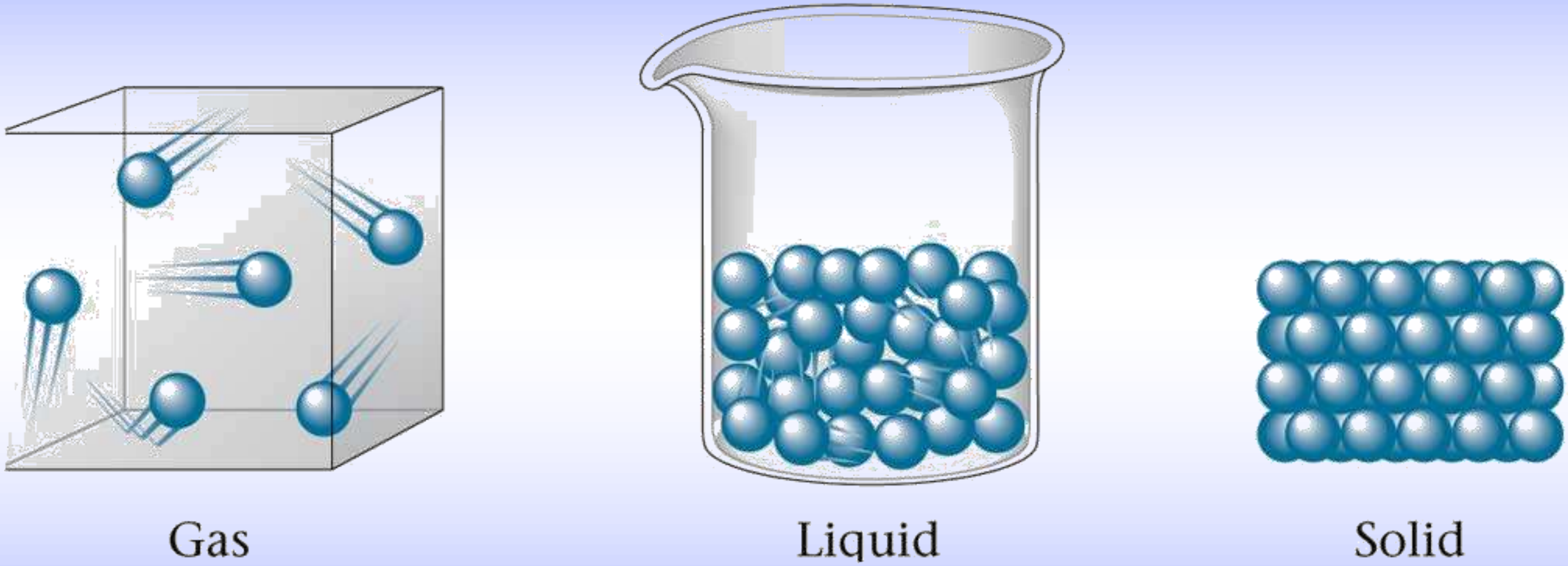
Heating Curve

- As heat added to solid, it first raises the temperature of the solid to the melting point
- Then added heat goes into melting the solid
  - Temperature stays at the melting point
  - Heat of Fusion
- As more heat added it raises the temperature of the liquid to the boiling point
- Then added heat goes into boiling the liquid
  - Temperature stays at the boiling point
  - Heat of Vaporization
- As more heat added it raises the temperature of the gas

Structures of the States of Matter

- In **solids,** the molecules have no translational freedom, they are held in place by strong **intermolecular** attractive forces
  - may only vibrate
- In **liquids,** the molecules have some translational freedom, but not enough to escape their attraction for neighboring molecules
  - they can slide past one another, rotate as well as vibrate
- In **gases,** the molecules have “complete” freedom from each other

Representations of the gas, liquid, and solid states.



Energy Requirements for State Changes

- In order to change a liquid to a gas, must supply the energy required to overcome the all the intermolecular attractions
  - Not break bonds (intramolecular forces)
- The energy required to boil 1 mole of a liquid is called the **Heat of Vaporization**
  - $\Delta H_{vaporization}$ = 40.6 kJ/mol for water at 100°C

Energy Requirements for State Changes

- In order to change a solid to a liquid must supply the energy required to overcome the some of the intermolecular attractions
- The energy required to melt 1 mole of a solid is called the **Heat of Fusion**
  - ΔH_{fusion} = 6.02 kJ/mol for ice at 0°C

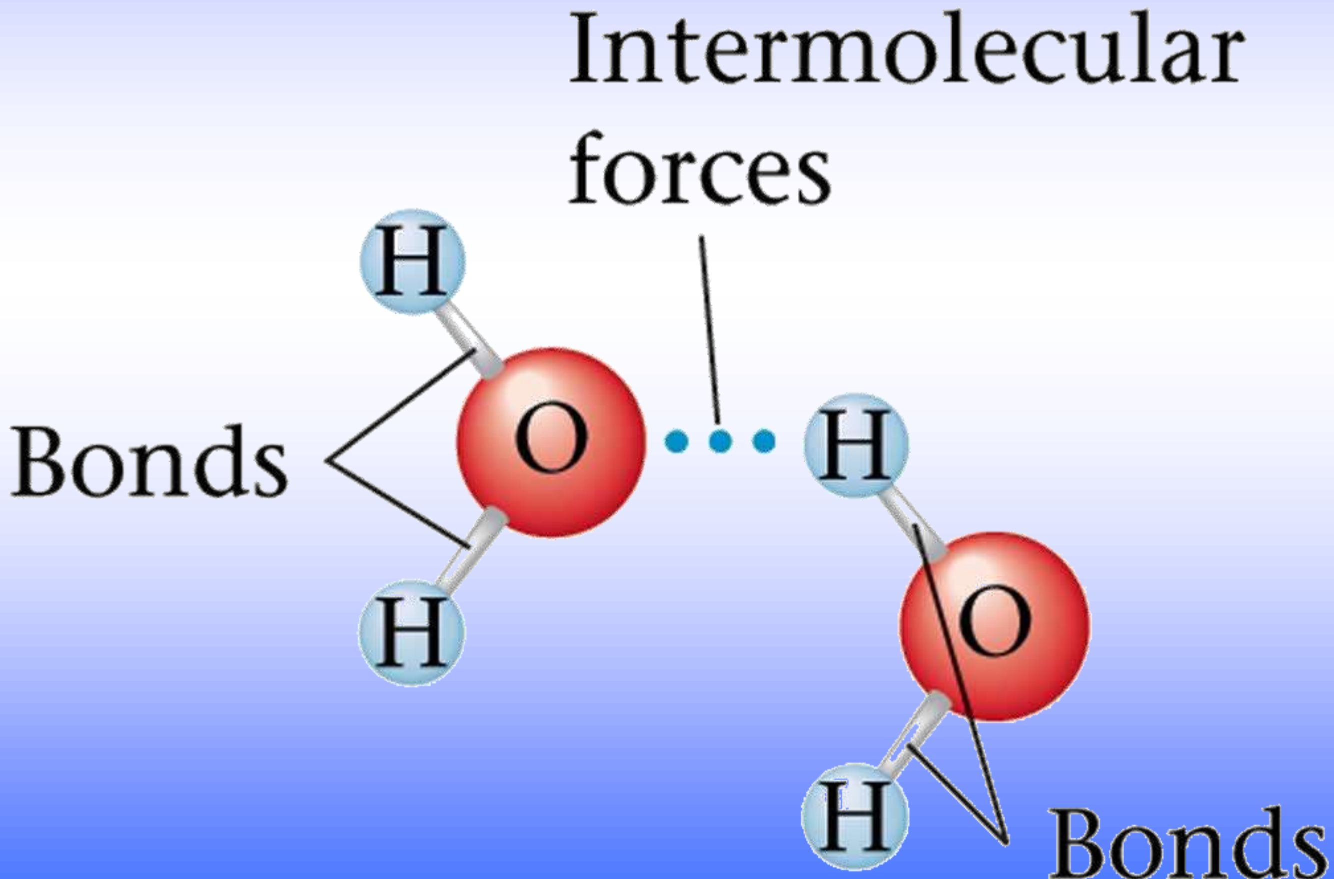
Why do Molecules Attract Each Other?

- Intermolecular attractions are due to attractive forces between opposite charges
- + Ion to - ion
- + End of polar molecule to - end of polar molecule
  - H-bonding especially strong
- **Larger the charge = Stronger attraction**
- Even non-polar molecules have attractions due to opposite charges
  - London Dispersion Forces

Intermolecular Attraction

- Much weaker than ionic or covalent bonds
  - Become weaker as distance between molecules increases
- Dipole-to-Dipole intermolecular attraction due to molecular polarity
- Hydrogen Bond is a special type of very strong dipole-to-dipole intermolecular attraction
  - Water has very strong H-bonds
- London Dispersion Forces are intermolecular attractions between non-polar molecules

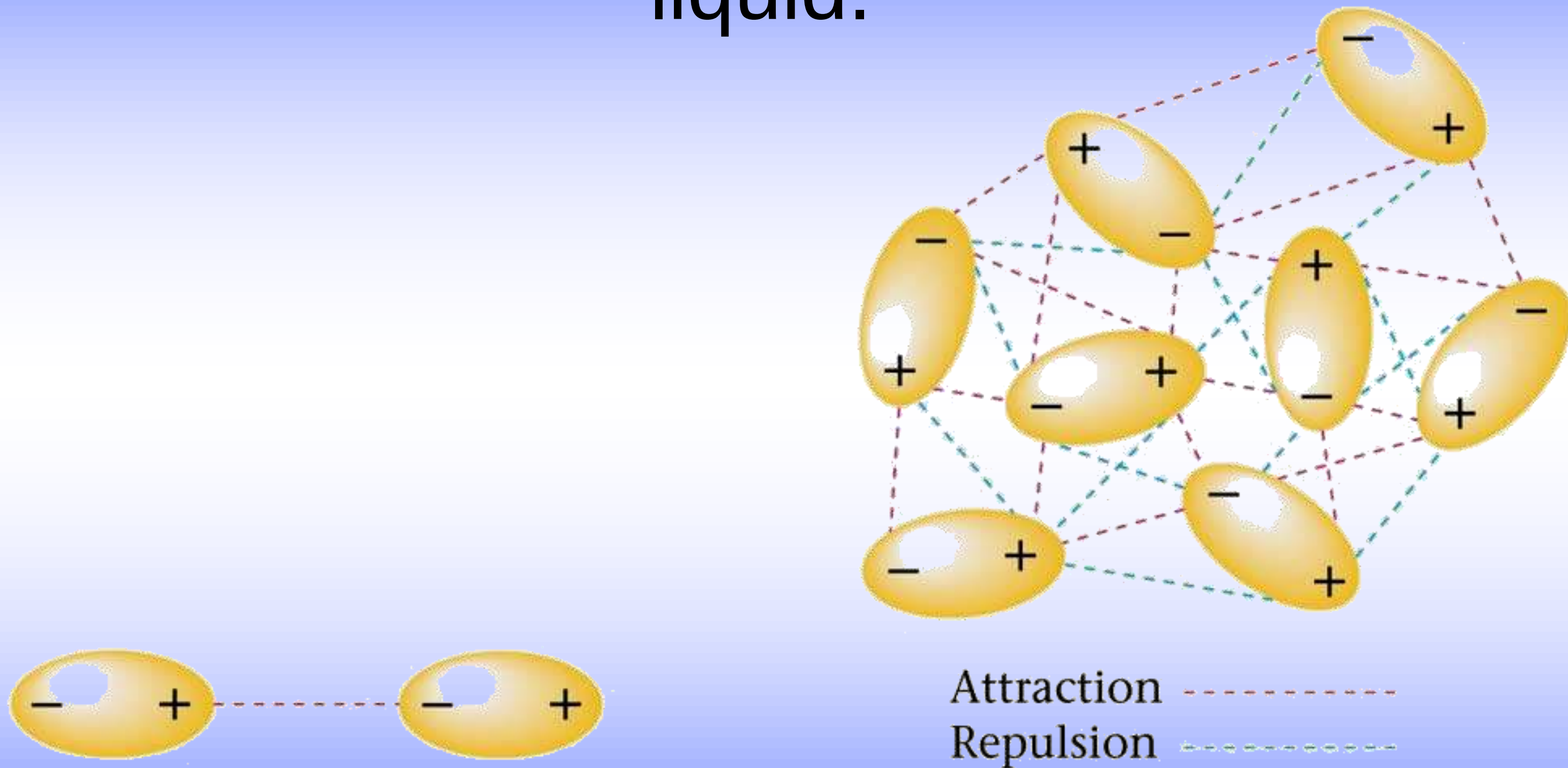
Intramolecular (bonding) forces exist between the atoms in a molecule and hold the molecule together. Intermolecular forces exist between molecules.



Dipole-to-Dipole Attractions

- Size of permanent dipole depends on the bonded atoms and shapes of molecules
- Individually, dipole-to-dipole attraction is stronger than induced dipole-to-induced dipole attraction
  - But for larger molecules the London Dispersion forces become more important for predicting the physical properties

(a) The interaction of two polar molecules. (b) The interaction of many dipoles in a liquid.

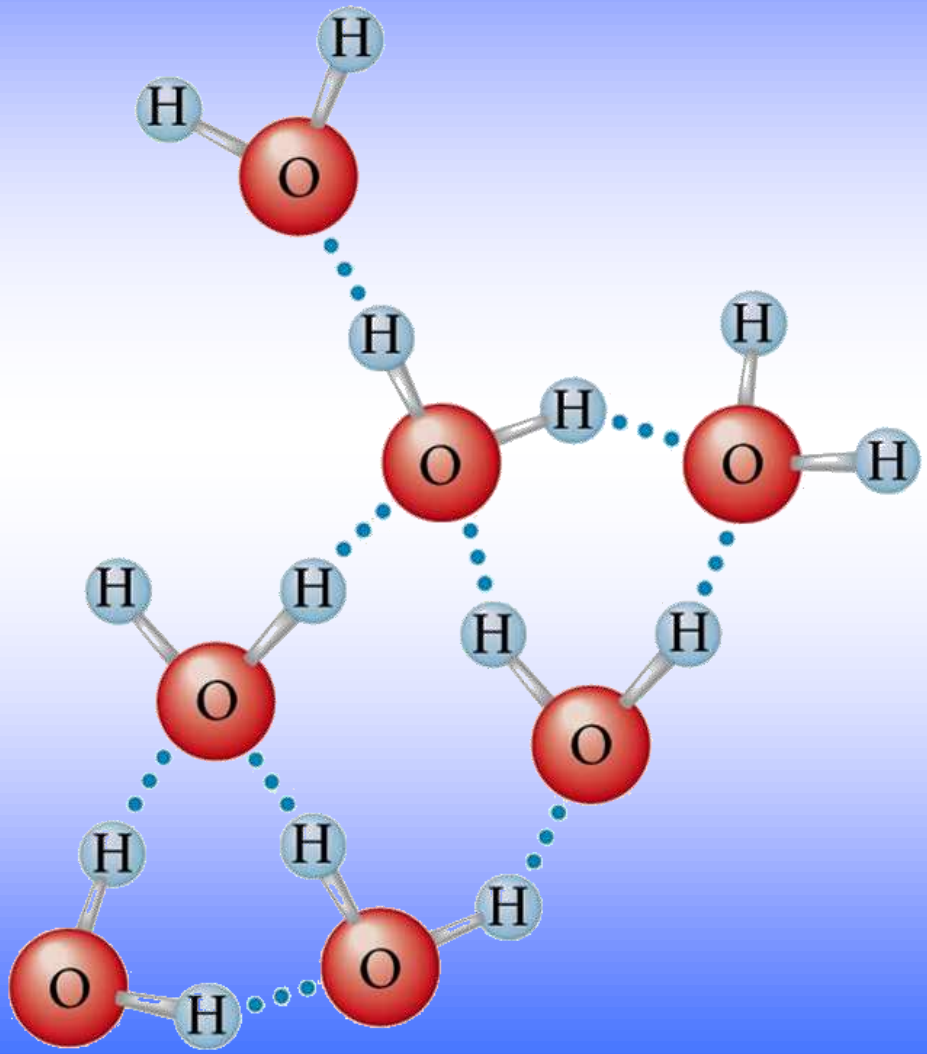


(a) (b)

Hydrogen Bonding

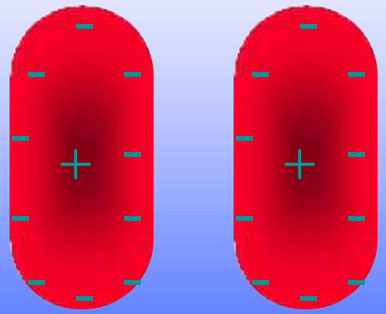
- Molecules that have -**OH** or -**NH** groups have particularly large intermolecular attractions
  - Also the HF molecule
  - unusually high melting and boiling points
  - unusually high solubility in water
- As electrons are pulled away from H by an electronegative atom, what is left is an unshielded proton that will strongly attract neighboring electrons

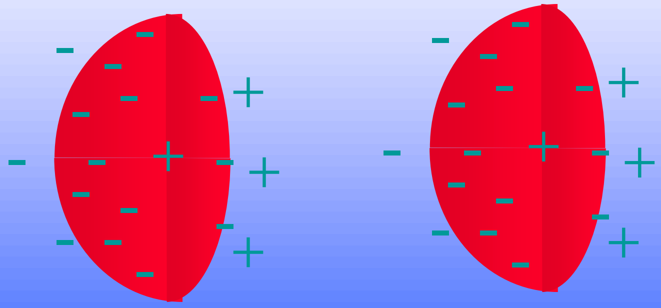
Hydrogen bonding among water molecules.



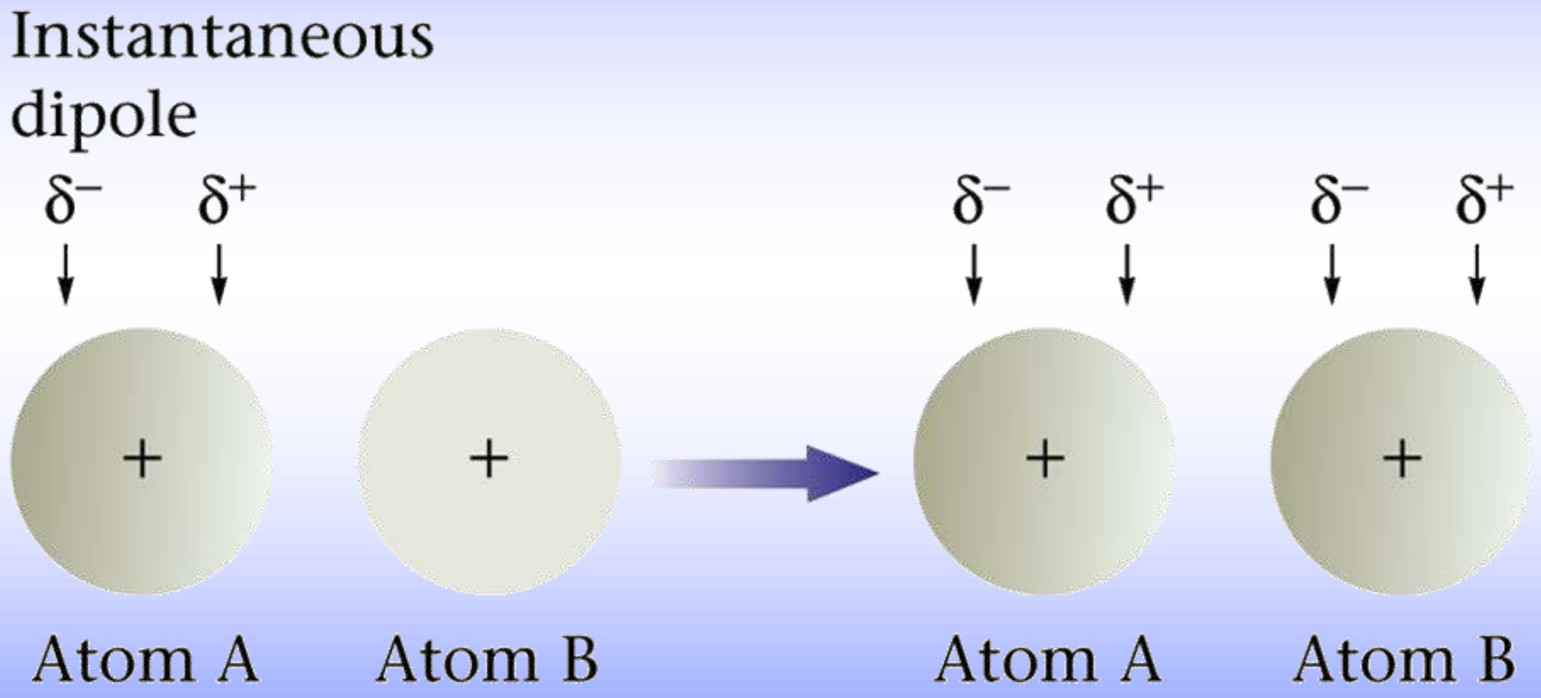
London Dispersion Forces

- Also Known As Induced Dipoles
- Caused by electrons on one molecule distorting the electron cloud on another
- All molecules have them
- Temporary
- Size of the London Dispersion Force depends on the number of electrons and shapes of molecules
  - the larger the molar mass, the larger the induced dipole

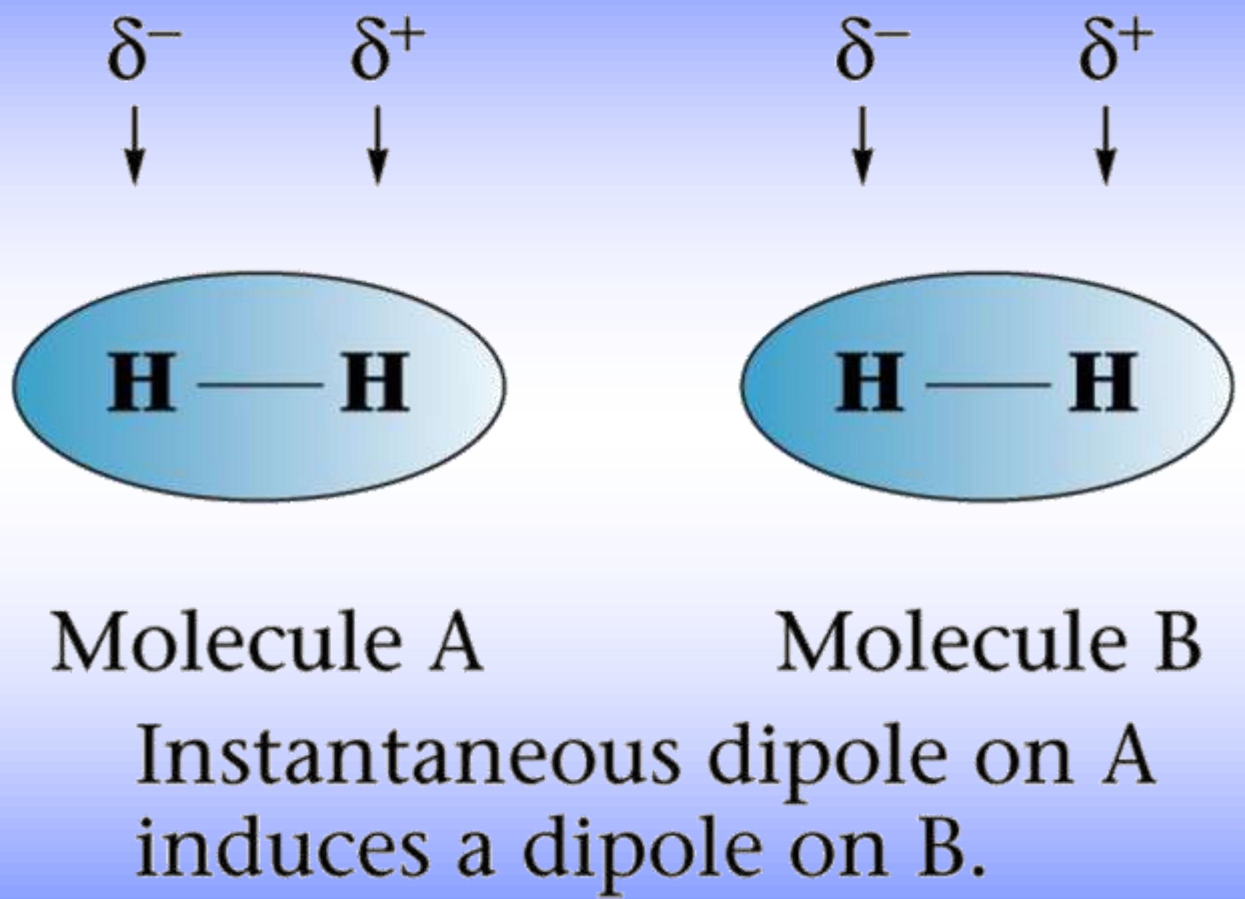




The atom on the left develops an instantaneous dipole when more electrons happen to congregate on the left than on the right.



Nonpolar molecules also interact by developing instantaneous dipoles.



Attractive Forces and Properties

- Larger attractive forces between molecules in pure substance means
  - higher boiling point
  - higher melting point (though also depends on crystal packing)
- Like dissolves Like
  - Polar molecules dissolve in polar solvents
    - Water, alcohol
    - Molecules with O or N higher solubility in H_2O due to H-bonding with H_2O
  - Non-polar molecules dissolve in non-polar solvents
    - Oils and gasoline

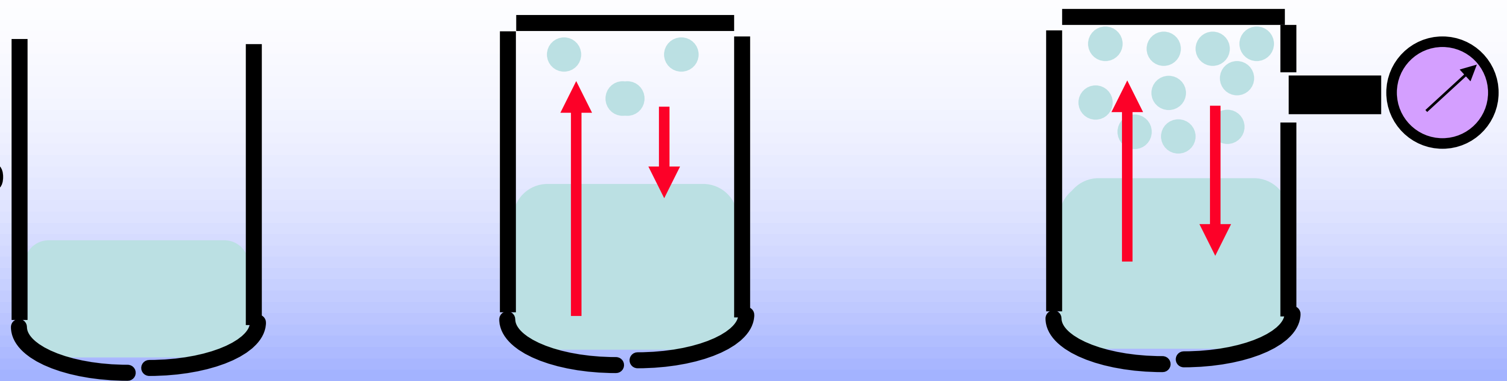
Evaporation

- Also Known As Vaporization
- Requires overcoming intermolecular attractions
- Heat of Vaporization is amount of energy needed to evaporate 1 mole of liquid
- Condensation the reverse
- In a closed container, eventually the rate of evaporation and condensation are equal
  - Equilibrium
  - In open system, evaporation continues until all liquid evaporated

Vapor Pressure

- Pressure exerted by a vapor in equilibrium with a liquid
  - Or solid
- Increases with temperature
- Larger intermolecular forces = Lower Vapor Pressure
- **Liquid boils when its Vapor Pressure = Atmospheric Pressure**
  - Normal boiling point
  - Raising external pressure raises boiling point, & visa versa

Equilibrium



Liquid just poured into open container, little vapor

Evaporation faster than Condensation

Evaporation as fast as Condensation

Solids

- amorphous solids
  - show no definite structure
    - therefore strengths of intermolecular forces vary over the structure
    - glass, plastic, rubber
  - tend to soften and melt over a temperature range
- crystalline solids
  - orderly, repeating, 3-dimensional pattern
  - pattern = **crystal lattice**
  - melt at one specific temperature

Crystalline Solids

Metallic	Molecular	Ionic	Atomic Networks
malleable & ductile	brittle & weak, or soft & waxy solids	hard & brittle	very hard
Usually high MP	MP < 300°C	MP > 300°C	MP > 1000°C
High BP	Low BP	Very high BP	Very high BP
High ΔH_{vap}	Low ΔH_{vap}, ΔH_{fusion}	High ΔH_{vap}, ΔH_{fusion}	Very high ΔH_{vap}, ΔH_{fusion}
high density	low density	medium density	medium density
good conductor	insulator	good electrical conductor when molten or dissolved in water	very insulating very unreactive
soluble in other metals	solubility varies	often soluble in water	dissolve in very few things

Types of Crystalline Solids

- Type of solid depends on type of particle that makes it up
- Properties of crystalline solid depend on the forces of attraction between the particles
- Ionic solids are made up of cations and anions arranged in a pattern that maximizes the interaction between ions
  - Strong forces between ions
  - Attractions are for all surrounding ions

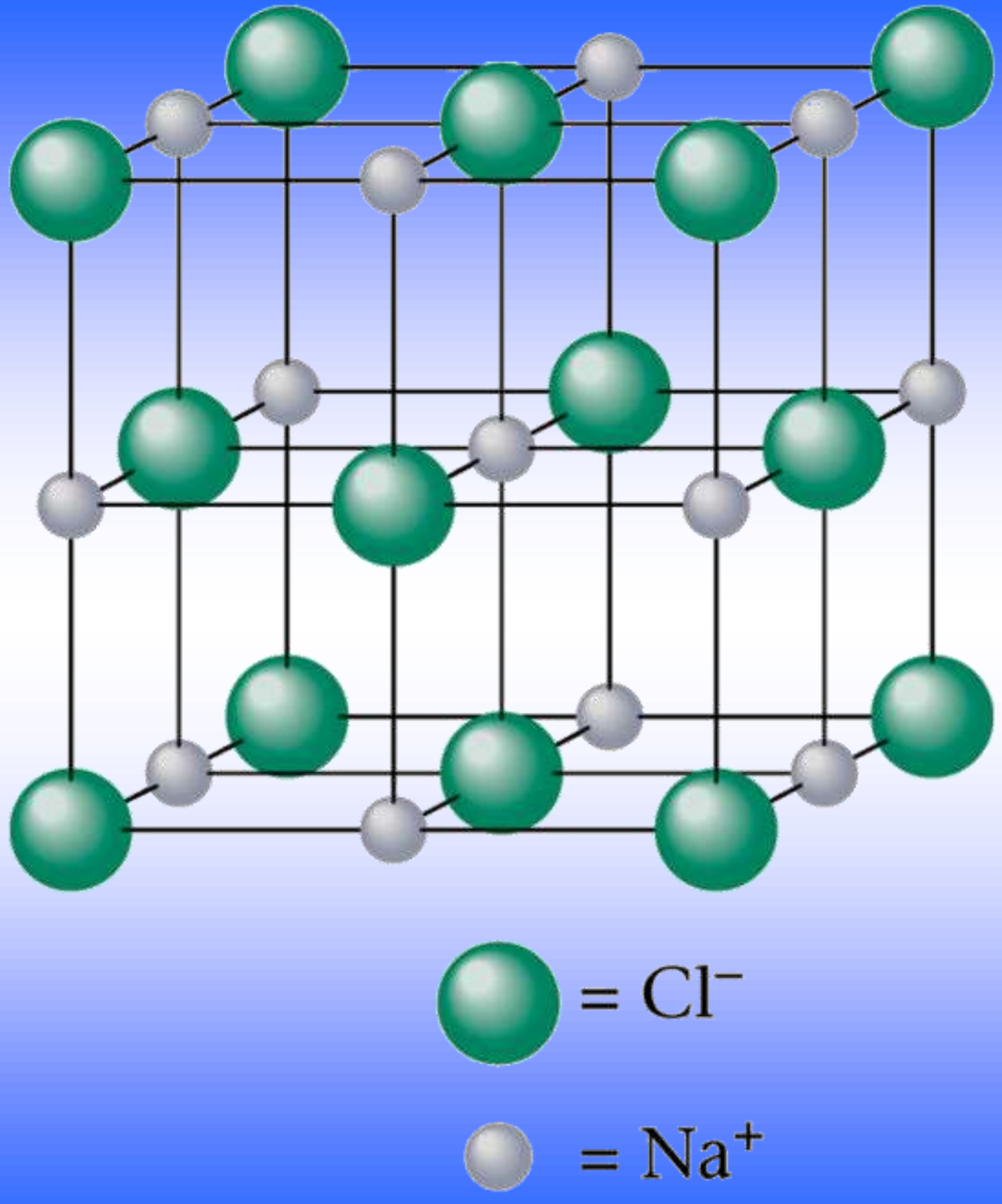
Types of Crystalline Solids

- Molecular Solids are made of molecules arranged in a pattern that maximizes the attractive forces between the molecules
  - Weak intermolecular attractive forces
- Atomic Solids of Noble Gases behave like molecular solids
  - London Dispersion Forces

Types of Crystalline Solids

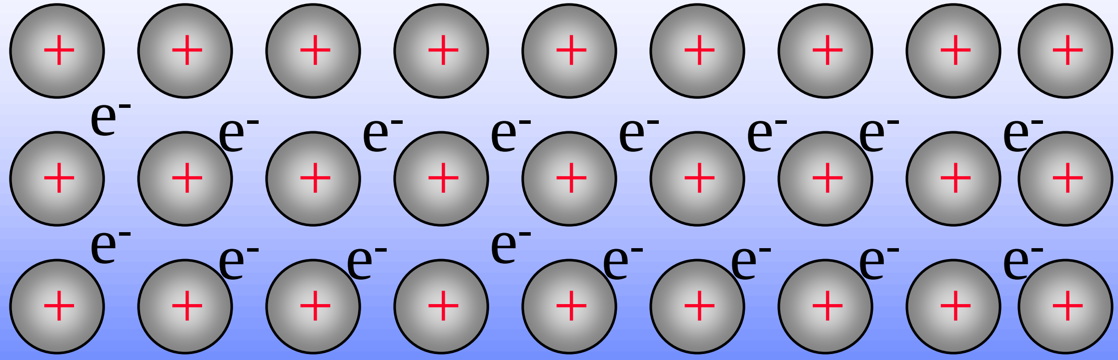
- Atomic Solids of the Non-metals and Metalloids in Group 4A generally have a network of atoms that are covalently bonded together
- Arrangement depends on VSEPR shape
  - May be 3-dimensional network (diamond), 2-dimensional network (graphite) or 1-dimensional network (asbestos)
  - Attraction between atoms is directional

The regular arrangement of sodium and chloride ions in sodium chloride, a crystalline solid.



Types of Crystalline Solids

- Atomic Solids that are made of metal atoms
  - metal atoms release their valence electrons
  - metal cations fixed in a "sea" of mobile electrons
  - Leads to strong attractions that are non-directional



Several crystalline solids: quartz, SiO_2; rock salt, $NaCl$; and iron pyrite, FeS_2.

Metallic Bonding & Properties

- ***Electrical Conductors*** because of mobile electrons
  - electrical current is moving electrons
- ***Thermal Conductors*** because of mobile electrons
  - small, light, mobile electrons can transfer kinetic energy quicker than larger particles
- ***Malleable & Ductile*** because metal ions can slide past each other in crystal
  - attractive forces are adjustable because electrons are mobile

Metallic Bonding & Properties

- ***Luster*** due to the mobile electrons on the surface absorbing the outside light and then emitting it at the same frequency
- ***Melting Points*** are relatively high due to strong attractions between metal ions and electrons
  - Stronger ion charge = Larger attraction
  - Smaller ion size = Larger attraction

Metallic Solutions

- **Alloys** are mixtures of elements that show metallic properties
- **Substitutional Alloys** have some host metal atoms replaced by metal atoms of similar size
  - **Brass**
- **Interstitial Alloys** have small atoms occupying some of the holes in the crystal lattice of the host metal
  - **Steel**