10.1 - 10.2 - Intermolecular Forces & Liquids

Overview

In Unit 7, we will look at interactions between particles in the gaseous, liquid, and solid state.

Firstly, this involves looking at forces between particles (which have been previously touched upon).

We will be looking at properties of the **condensed state**, either at solids or liquids, where particles are in close contact (as opposed to gases).

We will proceed to bulk properties: for example how solutions are made, and the properties that those solutions have.

Intermolecular Forces

We've seen previously how polar bonds can give rise to polar molecules. I**ntermolecular forces**, acting *between* particles (as opposed to chemical bonds, which are between atoms in a particle), give rise to many physical properties, such as melting and boiling points.

To review:
A. dipole-dipole: (see 8.1 - 8.4, and 8.13 in your Yellow Booklet) interactions are attractions between oppositely charged ends of polar molecules.
These are about 1% as strong as covalent or ionic bonds.

B. hydrogen bond: relatively strong dipole-dipole interaction where one pole is hydrogen, and the other is nitrogen, oxygen, or fluorine.

London Dispersion Forces

Even molecules without dipole moments exert forces on each other.

London dispersion forces: the forces of attraction that exist between noble gas atoms (else: hows could they condense and freeze at low enough temperatures?), and nonpolar molecules arise from an *instantaneous dipole* that arises from a asymmetric distribution of electrons in the atom.

We usually envision electron clouds as uniform, but indeed as the electrons move, they will briefly be more dense on one side of an atom, which can induce (cause to form) a similar distribution of electrons on a neighboring atom, and produce a fleeting dipole moment which attracts the atoms to each other.

Properties Driven by Forces

Melting/Freezing Point: temperature at which a substance undergoes a phase change from solid to liquid, or liquid to solid.

Boiling/Condensation Point: temperature at which a substance undergoes a phase change from liquid to gas, or gas to liquid.

Vapor: gaseous state of matter at a temperature at which a substance *can* also be a solid or liquid. Example: at room temperature, water is liquid, but there is water vapor in the air.

Properties Driven by Forces

Phase change temperatures are driven by two main factors, intermolecular forces and molecular weight.

The more attraction a particle has to its neighbors, the more energy requires to reach the next state of matter.

Also, the more massive it is, the more energy is required to make particles start to separate. That is due in part to increased London forces (with more electrons), as well as kinetic energy effects (Kinetic Energy = $1/2\ mv^2$, and kinetic energy of particles drives phase changes).

The Liquid State

Intermolecular forces between liquid particles give rise to important properties.

When water is dropped onto plastic or glass, it will naturally form spherical beads. Hydrogen bonding between water molecules pulls them together, and they the shape that minimizes surface area is a sphere.

Surface tension: the resistance a liquid has to an increase in surface area. Essentially, the molecules at the surface pull on each other, and resist the stretching action than an applied force delivers.

Capillary action: spontaneous rising of liquid in a narrow tube, brought about by the adhesion of polar molecules to the surface.
Example: water at the edge of glass tube will be higher than at the middle of the column.

The Liquid State (Continued)

Viscosity: a measure of a liquid's resistance to flow. Molecules with large intermolecular forces will flow less easily than those with weaker forces.
Ex: water flows more easily than honey, as large sugar molecules attract each other, slowing down mass transfer.
BUT: alcohol flows easier than water, as the hydrogen bonding in alcohol, while present, is less than in water.

Vapor Pressure: pressure that an evaporating compound exerts on its environment at the surface of its liquid state.
Molecules with strong intermolecular forces have lower vapor pressure than those with weak ones.

Boiling Point Examples

For the following sets of comparable compounds, list them in order of increasing boiling points. Explain your reasoning for your order - which forces contribute to the placement you chose?

1. The noble gases: Xe, He, Ar, Ne, Kr
2. Hydroacids: HBr, HI, HF, HCl
3. Group 4 hydrides: SnH_4, SiH_4, CH_4, GeH_4.
4. Group 5 hydrides: NH_3, PH_3, SbH_3, AsH_3.
5. Group 6 hydrides: H_2Se, H_2S, H_2O, H_2Te.

Boiling Point Answers

1. He, Ne, Ar, Kr, Xe. London forces are all that act on the noble gases. As more electrons are added, more London force exists. Also, increased mass requires more energy to overcome.
2. HCl, HBr, HI, HF. HF has hydrogen bonding, and will boil last due to this. The others rank according to increased mass and London forces: dipole-dipole forces don't play a major role.
3. CH_4, SiH_4, GeH_4, SnH_4. These are directly linked to increased mass, and London forces.
4. PH_3, AsH_3, NH_3, SbH_3. Hydrogen bonds make NH_3 boil off after PH_3 and AsH_3, but are not strong enough to place it after SbH_3: most massive.
5. H_2S, H_2Se, H_2Te, H_2O. H_2O has hydrogen bonding, and boils last due to this. The others rank according to increased mass and London forces.

Other Terms

Molar Heat (Enthalpy) of Fusion: a measure of energy released when one mole of a liquid freezes at its freezing point. Note that the same amount of energy is required to melt one mole of a solid into a liquid at its melting point.
As you might suspect, the greater the intermolecular attraction, the more energy is required to disrupt bonds.

Molar Heat (Enthalpy) of Vaporization: energy required to evaporate one mole of a liquid at its boiling point. The same amount of energy is released during the condensation of one mole.
The greater the attraction, the larger this value.

Homework

Preview 10.3 - 10.5

10.1 - 10.2 Problems in your Booklet
Due: Next Class