Chapter 10 Notes

As taught by Ms. Tracey Pannapara, 2017-18 term
Chemistry Lecture Notes
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1 Gases

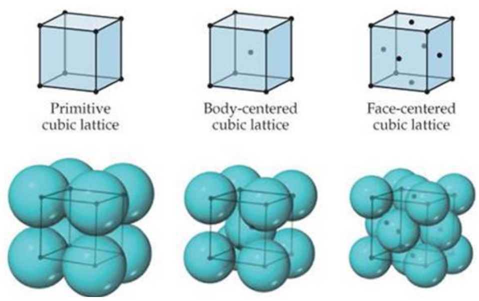
1. Kinetic-molecular theory makes five assumptions for ideal gases:
   (a) Particles have no volume compared to distance between them
   (b) Collisions are elastic
   (c) Particles are continually moving randomly and rapidly
   (d) No forces of attraction between particles
   (e) Temperature depends on average kinetic energy
2. KMT explains a lot of the properties of gases:
   - Expansion: particles move rapidly in all directions (1c) with no attractive forces (1d)
   - Fluidity: particles have very low attractive forces (1d)
   - Low density: particles far apart (1a)
   - Compressibility: particles very far apart (1a)
   - Diffusion/Effusion: constant motion (1c)
3. Rate of diffusion/effusion proportional to velocity, so molecules of lower mass will effuse faster (given same temperature, so same average kinetic energy)
4. No “real gas” behaves exactly like an ideal gas, but we can get close by putting a nonpolar gas at high temperature and low pressure

2 Liquids

1. Liquids are also in motion, but have intermolecular forces that give them less mobility than gases (still fluids)
2. Particles closer together: most liquids are hundreds of times denser than their gases
3. Liquids are relatively incompressible
4. Liquids diffuse like gases, but much more slowly (slower motion and stronger attraction between particles)
5. Surface tension: intermolecular forces pull adjacent parts of a liquid’s surface together, tending to minimize surface area (e.g. in water, surface molecules form H-bonds with interior molecules, giving droplets an optimal spherical shape)
6. Capillary action: attraction between liquid and solid surface, causing liquid to be drawn up along surface until balanced by gravity; responsible for plant water transport and meniscus formation
7. Evaporation: a form of vaporization in which surface particles with above average kinetic energies overcome IMFs and escape into the gaseous state

3 Solids

1. Particles more closely packed with stronger IMFs, causing only vibration of particles
2. Properties of solids:
   (a) Definite shape and volume: volume changes only slightly with a change in temperature or pressure, since there is little empty space available
   (b) High density and incompressibility: solids are typically slightly denser than liquids and much denser than gases; for practical purposes, solids are basically incompressible
   (c) Low rate of diffusion: diffusion does happen, but is is extremely slow
3. Crystalline solids: particles arranged in pattern
   (a) Definite melting point: kinetic energies of particles overcome attractive forces
   (b) Unit cell: know the three different types (replace "primitive" with "simple"):



   (c) Different types of crystalline solids:
      i. **Ionic:** Positive and negative ions; brittle and have pretty high melting points
      ii. **Covalent Network:** Very strong covalent bonds through a network (basically one giant molecule); very high melting points (e.g. diamond, quartz)
      iii. **Metallic:** Sea of electron stuff; melting point varies
      iv. **Covalent Molecular:** Covalently bonded molecules held together by intermolecular forces; could be polar or nonpolar; low melting points
4. Amorphous solids: particles arranged randomly
   (a) No definite melting point: can flow over a range of temperatures due to random arrangement (also known as supercooled liquids)
   (b) Examples: glasses, plastics, semiconductors

4 Changes of State

1. In a closed system, the rates of condensation and vaporization reach equilibrium
2. Equilibrium vapor pressure is proportional to number of molecules in the vapor phase, so higher temperature causes higher equilibrium vapor pressure; also, stronger IMFs lead to fewer particles in the vapor phase, lowering the equilibrium vapor pressure
3. Boiling
   (a) Boiling point is when equilibrium vapor pressure equals atmospheric pressure (so lower atmospheric pressure causes lower boiling temperature)
   (b) Energy must be continually added to keep a substance boiling, but that energy (the molar enthalpy of vaporization), goes toward breaking attractive forces instead of raising temperature

(c) Molar enthalpy of vaporization (ΔH_v) has a negative sign for condensation (take away energy)

(d) Higher boiling point means that one can heat up liquids more, causing faster cooking time

4. Freezing

   (a) In a pure crystalline substance, this happens at a constant temperature

   (b) At equilibrium, melting and freezing happen at the same rate: solid + energy $\rightleftharpoons$ liquid

   (c) If you disturb the equilibrium by adding heat, the forward reaction of solid $\rightarrow$ liquid will be favored

   (d) Molar enthalpy of fusion (ΔH_f), just like ΔH_v, depends on strength of attraction between particles

   (e) Sublimation and deposition equation is basically the same thing

5. Phase diagram: triple point and critical point

6. Overview of phase changes:

Process	Change of State
melting	solid + energy $\longrightarrow$ liquid
sublimation	solid + energy $\longrightarrow$ gas
freezing	liquid $\longrightarrow$ solid + energy
vaporization	liquid + energy $\longrightarrow$ gas
condensation	gas $\longrightarrow$ liquid + energy
deposition	gas $\longrightarrow$ solid + energy