### Chapter 11 part 2: Properties of Liquids

**Read:** BLB 5.5; 11.4  
**HW:** BLB 5:48, 49, 51; 11:33, 37, 39  
Packet 11:5-10

**Know:**  
- viscosity, surface tension  
- cohesive & adhesive forces  
- phase changes  
- heat capacity  
- calorimetry

<table>
<thead>
<tr>
<th>Functional Group</th>
<th>boiling point</th>
<th>Structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrocarbon</td>
<td>36°C</td>
<td>CH₃CH₂CH₂CH₂—CH₃</td>
</tr>
<tr>
<td>aldehyde</td>
<td>75°C</td>
<td>CH₃CH₂CH₂—C=O</td>
</tr>
<tr>
<td>ketone</td>
<td>79°C</td>
<td>CH₃CH=O</td>
</tr>
<tr>
<td>amine</td>
<td>78°C</td>
<td>CH₃CH₂CH₂CH₂—NH₂</td>
</tr>
<tr>
<td>ether</td>
<td>34°C</td>
<td>CH₃CH₂—O—CH₂CH₃</td>
</tr>
<tr>
<td>ester</td>
<td>57°C</td>
<td>CH₃—C=O</td>
</tr>
<tr>
<td>alcohol</td>
<td>117°C</td>
<td>CH₃CH₂CH₂CH₂—OH</td>
</tr>
<tr>
<td>carboxylic acid</td>
<td>141°C</td>
<td>CH₃CH₂—C=O</td>
</tr>
</tbody>
</table>

**Structure Affects Function**

- boiling point
- functional group
Manifestations of IM forces

- **cohesive forces**: forces between similar molecules
  
  **for example:**
  - viscosity: resistance to flow
  - surface tension: E needed to increase surface area
  - BP (MP)
  - $\Delta H_{\text{vap}}$ ($\Delta H_{\text{sub}}$)

- **adhesive forces**: forces between a substance & the surface
  
  **for example:**
  - water–glass interactions
  - water–oil interactions
  - meniscus formation
  - capillary action
  - glue
  - non-stick surfaces

Properties of Liquids

- Intermolecular interactions are favorable (heat is required to break them)
- The more interactions, the better.
- Surface molecules have fewer interactions.
- **Energy is minimized by minimizing the surface area.**

$\Rightarrow$ What shape gives a water drop the most favorable Energy?

**Meniscus Shapes**
- Cohesive Forces
- Adhesive Forces

**Capillary Action**
  - Plants get water, nutrients
Kinetic Molecular Description of Liquids and Solids

**Gas:**
- Kinetic energy $\gg$ intermolecular forces

**Liquid:**
- Kinetic energy $\approx$ intermolecular forces

**Solid:**
- Kinetic energy $\ll$ intermolecular forces

Kinetic Energy $\propto T$

Heating: $T \uparrow$, KE $\uparrow$

**Solid $\rightarrow$ Liquid $\rightarrow$ Gas**

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**Energy of phase changes**

**Endothermic**
- Requires energy to disrupt intermolecular forces.
- Which processes are endothermic?

**Exothermic**
- Energy is released when intermolecular interactions are formed.
- Which processes are exothermic?

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Endothermic processes:
- Sublimation
- Vaporization
- Deposition

Exothermic processes:
- Freezing
- Melting
- Condensation
Heat Capacity

Heat Capacity (C):

Usually give C for a specified amount of pure substance
Example: liquid H₂O
1 g Cₜₕ = s = 4.184 J/g-K → specific heat
1 mol Cₘₖₜ = 75.2 J/mole-K → molar heat capacity

Specific heat is different for each phase
Specific heat is different per gram than per mole

The specific heat for water is very high compared to other substances; 5 times greater than Al(s).

Why is this important?

CALORIMETRY
Experimental measure of heat flow

\[ q_{\text{soln}} = C_s \ m \ \Delta T \]

\[ q = \text{heat flow (}\Delta H\text{ at } P_{\text{const}}) \]

\[ C = \text{specific heat} \]

\[ m = \text{mass} \]

\[ \Delta T = T_{\text{final}} - T_{\text{initial}} \]

\[ \Delta H = \text{Heat change} = q_{\text{final}} - q_{\text{initial}} \]

= amount of heat given off (−) or absorbed (+) when a change occurs

Examples:
\[ \Delta H_{\text{fusion}} = \text{amount of heat needed to} \]

\[ \Delta H_{\text{vap}} = \text{amount of heat needed to} \]

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**Example:**

\[
\text{NaOH(s)} \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)
\]

When 9.55 g solid NaOH is dissolved in 100.0 g of water in a coffee cup calorimeter, the temperature rises from 23.6 °C to 47.4 °C. Calculate \(\Delta H\) in kJ/mol for the solution process. The specific heat of water is 4.184 kJ/g-°C

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**Bomb Calorimetry**

\[
q_{rxn} = (-C_{cal})(\Delta T)
\]

Where \(C_{cal}\) =
Heating Curve for 1 mole of Water

Two types of changes; Phase Changes and Heating a single Phase

1. Within a single phase (blue) changes are continuous

\[ q = n \ C_p \ \Delta T \]  
(p \Rightarrow \text{constant pressure})

T is increasing . . .

What happens to kinetic energy?
molecular motion?
separation between molecules?
molecular attractions?
amount of order?

Phase Changes

2. Between phases, (red) changes are abrupt, from one physical state to another

\[ q = n \ \Delta H_x \]  
(x = fusion or vaporization)

T is constant . . .

What happens to kinetic energy?
molecular motion?
total energy?
separation between molecules?
molecular attractions?
amount of order?
Calculation of Enthalpy Change

Measure the heat required for each segment; take the sum of the individual steps.

EXAMPLE:
2 moles of ice at –25 °C are heated to 125 °C.
How much energy is needed?

\[
\begin{align*}
C_p(\text{ice}) &= 37.6 \text{ J/mol·K} & \Delta H_{\text{fusion}} &= 6.02 \text{ kJ/mol} \\
C_p(\text{water}) &= 75.3 \text{ J/mol·K} & \Delta H_{\text{vapor}} &= 40.67 \text{ kJ/mol} \\
C_p(\text{steam}) &= 33.1 \text{ J/mol·K} \\
\end{align*}
\]

1. ice –25 °C → 0 °C

2. ice 0 °C → water 0 °C  \hspace{1cm} \text{(phase transition)}

3. water 0 °C → 100 °C

4. water 100 °C → steam 100 °C  \hspace{1cm} \text{(phase transition)}

5. steam 100 °C → 125 °C