3.4 INTERMOLECULAR FORCES

Grade 11 University Chemistry
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PREVIOUS KNOWLEDGE:

The Basic Units: Ionic vs. Covalent

- Ionic compounds form repeating formula units.
- Covalent compounds form distinct molecules.
- Consider adding to NaCl(s) vs. H₂O(s):

![Ionic vs. Covalent Compounds](image)

The Basic Units: Ionic vs. Covalent

- NaCl: [Na⁺][Cl⁻]
- H₂O: H₂O

The Basic Units: Innermolecular Bonding

1. Non-polar covalent
2. Polar covalent
3. Ionic

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Example</th>
<th>Partial Charges</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂</td>
<td>H:H</td>
<td>δ₀, δ₀</td>
</tr>
<tr>
<td>HCl</td>
<td>H:Cl</td>
<td>δ⁺, δ⁻</td>
</tr>
<tr>
<td>LiCl</td>
<td>[Li⁺][Cl⁻]</td>
<td>+, −</td>
</tr>
</tbody>
</table>

The greek symbol δ indicates “partial charge”.

Intramolecular Bonding

1. Non-polar covalent
2. Polar covalent
3. Ionic
4. Metallic

Vector Addition & Bond Polarity

- The magnitude of a vector relates to bond polarity.
- The longer the vector the greater the dipole within a polar covalent bond.
- The vector always points towards more electronegative atom.
- The lone pair of electrons is represented by the longest vector.

<table>
<thead>
<tr>
<th>Electro negativity</th>
<th>C</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>2.55</td>
<td>3.16</td>
</tr>
<tr>
<td>F</td>
<td>3.98</td>
<td>3.98</td>
</tr>
</tbody>
</table>

δ⁺C-δ⁻F = 3.98 - 2.55 = 1.43
δ⁺C-δ⁻Cl = 3.16 - 2.55 = 0.61

Polarity of Molecules

The polarity of a molecule depends on the polarity of its bonds and the shape of the molecule.

Molecules have different shapes based on their electron arrangements.
An intermolecular force is an attraction that occurs between molecules. Increasing energy allows molecules to overcome these intermolecular forces. Intramolecular forces (ionic and covalent bonds) are significantly stronger than intramolecular forces.

The strength of intermolecular forces depends on the following physical properties of molecular compounds:
- Physical state of matter
- Melting and Boiling Point
- Surface Tension
- Hardness and Texture
- Solubility in Various Solvents
**Intermolecular Forces**

1. **Dipole-Dipole**
   - Dipole-dipole forces are attractive forces between the positive end of one polar molecule and the negative end of another polar molecule.
   - Dipole-dipole forces are much weaker than ionic or covalent bonds and have a significant effect only when the molecules involved are close together (touching or almost touching).
   - They are very strong intermolecular forces.

2. **London Dispersion**
   - When the vectors cancel each other, the molecule has the electron density distributed evenly throughout the molecule and has no permanent dipole. Ex. Boron trifluoride, carbon dioxide

3. **Hydrogen Bonding**
   - Hydrogen bonds are a stronger form of dipole-dipole interactions that occur between two molecules.
   - They are caused by highly electronegative atoms. They only occur between hydrogen and oxygen, fluorine or nitrogen, and are the strongest intermolecular force.
   - The high electronegativities of F, O and N create highly polar bonds with hydrogen, which leads to strong bonding between hydrogen atoms on one molecule and the lone pairs of F, O or N atoms on adjacent molecules.
**Intermolecular Forces**

3. **Hydrogen Bonding**

- An ion-dipole force is an attractive force that results from the electrostatic attraction between an ion and a neutral molecule that has a dipole.

**Bond Type Dissociation Energy**

<table>
<thead>
<tr>
<th>BOND OR FORCE</th>
<th>DISSOCIATION ENERGY (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic Bond</td>
<td>&gt;600</td>
</tr>
<tr>
<td>Covalent Bond</td>
<td>74-500</td>
</tr>
<tr>
<td>Hydrogen Bonds</td>
<td>16-70</td>
</tr>
<tr>
<td>Dipole-Dipole Forces</td>
<td>2.0-8.0</td>
</tr>
<tr>
<td>London Dispersion Forces</td>
<td>&lt;4.0</td>
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</table>