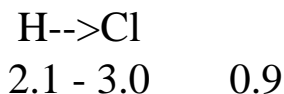


- 1  Properties of Liquids, Solids, & Solutions  
Chapters 10
- 2  10.1 Review Bond Polarity. Predict the direction of a dipole in a polar molecule  
[Readings 10.1 Problems 2 & 3]
- 3  10.2 Classify intermolecular bonds and predict relative properties of chemical substances.  
[Readings 10.1 - 10.3 ; Problems 4, 5, 36, & 38]
- 4  Intermolecular Bonds  
Intramolecular Bonds - chemical bonds between atoms or ions in a molecule or formula unit.  
Ionic and Covalent Bonds are examples  
Intermolecular Bonds - chemical bonds (attraction forces) between molecules or between molecules and ions.  
Intermolecular forces make it possible for the phases of matter to exist.  
The stronger intermolecular forces are, the greater the energy necessary to break them and the higher the melting and boiling points.

## 5 Intermolecular Forces & Polarity

Polar bonds (dipoles) are the result of differences in electronegativity between atoms.



The resultant force of polar bonds determines the polar nature of a molecule.

Compare  $\text{CCl}_4$  and  $\text{H}_2\text{O}$ .

## 6 Bonding Involving Permanent Dipoles

$\text{H}_2\text{O}$  - Polar Molecule

$\text{CCl}_4$  - Nonpolar molecule

## 7 Bonding Involving Permanent Dipoles

Two types of interaction involving dipoles:

Dipole - Ion

Dipole - Dipole

## 8 Hydrogen Bonding

Some dipole - dipole interactions are stronger.

If H is in a compound bonded to O, N, or F,

Then the resulting dipole has extra strength.

Why?

Two conditions:

atom must be highly electronegative

atom must be small

These bonds are extremely polar

9  **Hydrogen Bonding**

10  **Bonding Involving Induced Dipoles**

Induced dipoles cause London Dispersion Forces.

11  **Bonding Involving Induced Dipoles**







Induced dipoles cause London Dispersion Forces.

Larger molecules have larger LDF

12  **Strength of Intermolecular Forces**

<u>Type</u>	<u>Strength (kJ/mol)</u>
Ion - Dipole (Na <sup>+</sup> - water)	40-600
Dipole - Dipole (HCl) (including H-bonding)	5-25
Dipole - Induced Dipole (O <sub>2</sub> - water)	2-10
Induced Dipole - Induced Dipole (I <sub>2</sub> )	0.05-40

These forces are all relatively weak when compared to the strength of covalent and ionic bonds.

- 13  Normal BPs of Hydrogen Halides
- 14  10.3 Describe the structure and properties of liquids.  
[Readings 10.3 - 10.5; Problems]
- 15  Properties of Liquids
- Viscosity
  - Surface Tension
  - Both are due to the effect of strong intermolecular attractive forces
- 16  Evaporation
- Consider a sample of Benzene:  
 $C_6H_6(l) \rightarrow C_6H_6(g)$  @  $20^\circ C$   
Normal Boiling Point of  $C_6H_6$  is  $80^\circ C$   
Explain evaporation.
- 17  Evaporation
- $= 1/2 mu^2 = cT$
- = average energy
- u = average velocity  
depends on T
- 18  Evaporation
- $= 1/2 mu^2 = cT$
- = average energy

$u$  = average velocity  
depends on  $T$

19  **Evaporation**

$$= 1/2 m\bar{u}^2 = cT$$

= average energy

$u$  = average velocity  
depends on  $T$

20  **Vapor Pressure**

Evaporation

Condensation

Equilibrium

21  **Vapor Pressure**

Evaporation

Condensation

Equilibrium

22  **Vapor Pressure**

Evaporation

Condensation

Equilibrium

The partial pressure of the gaseous benzene at equilibrium is the vapor pressure.

23  **Vapor Pressure**

Evaporation

## Condensation

### Equilibrium

The partial pressure of the gaseous benzene at equilibrium is the vapor pressure.

## 24 Vapor Pressure and Temperature

Vapor pressure is temperature dependent.

As temperature increases,  $P_v$  increases.

## 25 Vapor Pressure and Temperature

Under certain circumstances vapor will form in any part of liquid and form a bubble.

$$P_{\text{above surface}} = P_v$$

### Boiling Point

## 26 Vapor Pressure and the BP

Under certain circumstances vapor will form in any part of liquid and form a bubble.

$$P_{\text{above surface}} = P_v$$

### Boiling Point

## 27 10.4 Interpret phase diagrams. [Readings 10.12; Problems 16]

## 28 Phase Diagram

## 29 Questions

How could one boil water at room temperature?

## 30 Questions

How could one boil water at room temperature?

How could one liquefy propane gas?

31  Questions

How could one boil water at room temperature?

How could one liquefy propane gas?

What determines how high is the boiling point?

32  10.5 Describe the structure and properties of solids.

[Readings 10.6-10.11; Problems]

33  10.6 Explain and calculate the energy changes associated with phase changes.  
[Readings 8.7 and 10.4; Problems 10.42, 48, 50, & 52]

34  Changes in Water Upon Being Heated from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Macroscopic:

35  Changes in Water Upon Being Heated from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Microscopic:

1. molecules in a rigid arrangement. Mostly vibrational

motion.

36  Changes in Water Upon Being Heated  
from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Microscopic:

2. molecules break free of rigid arrangement. T doesn't change.

37  Changes in Water Upon Being Heated  
from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Microscopic:

3. increased amounts of rotational and translational motion.

38  Changes in Water Upon Being Heated  
from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Microscopic:

4. molecules break completely free of one another. T doesn't change.

39  Changes in Water Upon Being Heated  
from  $-20^{\circ}\text{C}$  to  $110^{\circ}\text{C}$

Microscopic:

5. molecules spread apart. Increased motion (vib, rot, & trans).

40  Compare Molecules in Phases

41  Calculate Enthalpies of Phase Changes

<u>Process</u>	<u>Enthalpy</u>
• Fusion (Melting)	$H_{\text{fusion}}$
• Freezing	- $H_{\text{fusion}}$
• Vaporization	$H_{\text{vap}}$
• Condensation	- $H_{\text{vap}}$
• Sublimation	
• Deposition	

42 □ Sample Problem

- Calculate the energy required to heat 10.0 g of water from  $-10.0\text{ }^{\circ}\text{C}$  to  $110.0\text{ }^{\circ}\text{C}$ , given the following information for water.
- $H_{\text{fus}} = 334\text{ J/g}$
- $H_{\text{vap}} = 2257\text{ J/g}$
- $C\text{ (l)} = 4.18\text{ J/g}^{\circ}\text{C}$
- $C\text{ (s)} = 2.06\text{ J/g}^{\circ}\text{C}$
- $C\text{ (g)} = 2.03\text{ J/g}^{\circ}\text{C}$