

## 11.4 The Nature of Intermolecular Forces

### London Dispersion Forces – Stronger than you might think

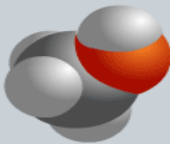
#### *Boiling Points of Simple Organic Molecules*

**Boiling Points of Simple Organic Compounds** Description

Alkyl Group	Functional Group
<input checked="" type="radio"/> Methyl	<input type="radio"/> Hydrogen
<input type="radio"/> Ethyl	<input checked="" type="radio"/> Hydroxyl
<input type="radio"/> Propyl	<input type="radio"/> Amine
<input type="radio"/> Isopropyl	<input type="radio"/> Chlorine
<input type="radio"/> Butyl	
<input type="radio"/> Pentyl	
<input type="radio"/> Hexyl	

Clear

Compound: methanol  
Boiling Point: 65 °C  
Molar Mass: 32 g/mol



Boiling Points, °C

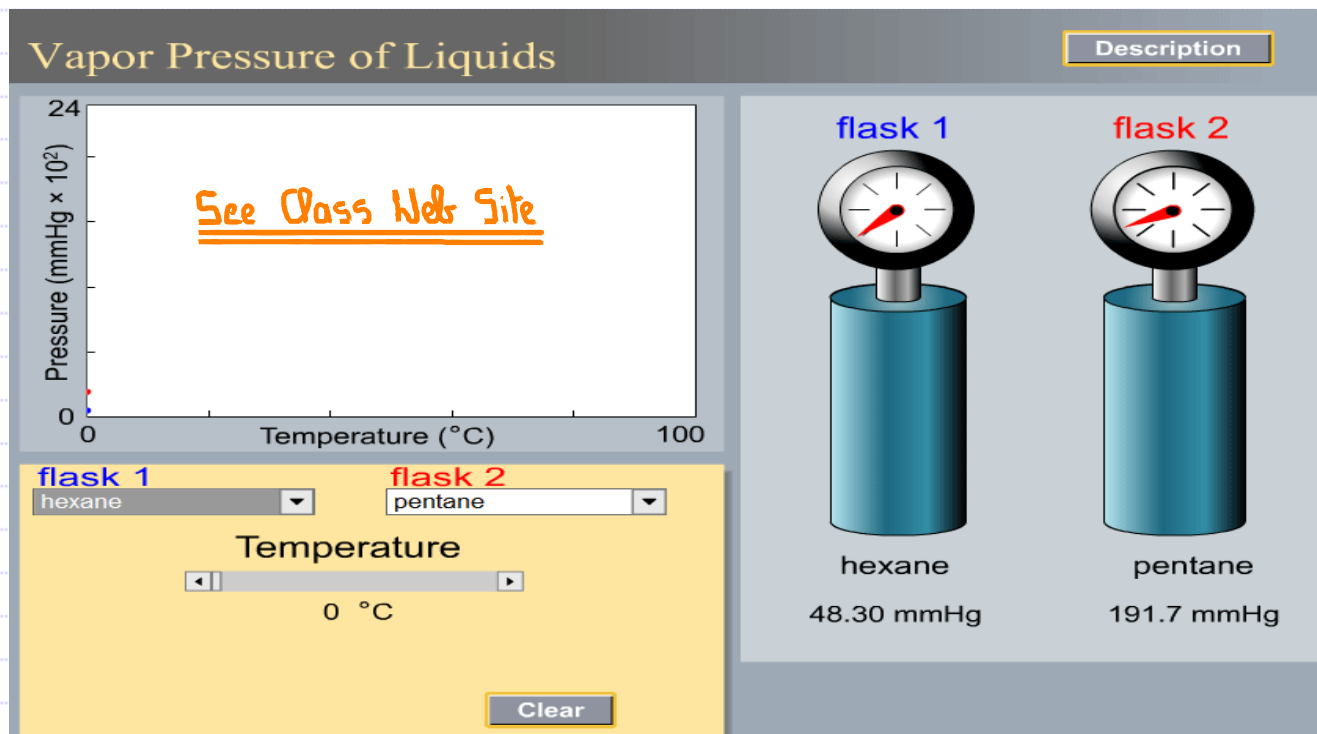
Molar Mass, g/mol

See Class Web Site

} O-H bond }  
N-H bond } Hydrogen Bond

Dipole - Dipole vs Induced Dipole - Induced Dipole  
(London Dispersion Forces)

## 11.2 Vapor Pressure



a) VP vs T

b) VP vs Molar Mass (non-polar)

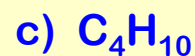
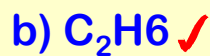
c) VP in polar molecules vs VP in nonpolar molecules.

## 11.2 Vapor Pressure Heat of Vaporization

→ The amount of heat required to convert a liquid to a gas:  $\Delta H_{\text{VAP}}^{\circ}$



Which of the following molecules would you expect to have the smallest  $\Delta H_{\text{vap}}^{\circ}$



→ Why?

Non-polar with a smaller Molar Mass than  $\text{C}_4\text{H}_{10}$ , which is also non-polar.

$\text{CH}_3\text{OH}$  is polar.

## 11.2 Vapor Pressure

### Relationship Between P, T, and $\Delta H^\circ_{\text{vap}}$ - Clausius-Clapeyron Equation

$$\ln P = \frac{-\Delta H^\circ_{\text{vap}}}{RT} + C$$

$\Delta H^\circ_{\text{vap}}$  = Heat of Vaporization.  
 $R$  : 8.314 J. mol<sup>-1</sup>. K<sup>-1</sup> (Ideal Gas Constant)

a) GRAPHICALLY:

Plot  $\ln P$  vs  $1/T$  : T must be in K

$$\text{Slope} = \frac{-\Delta H^\circ_{\text{vap}}}{R}$$

b) QUANTITATIVELY:

$$\ln P_1 = -\frac{\Delta H^\circ_{\text{vap}}}{RT_1} + C \quad : \quad \ln P_2 = -\frac{\Delta H^\circ_{\text{vap}}}{RT_2} + C$$

$$\ln P_2 - \ln P_1 = -\frac{\Delta H^\circ_{\text{vap}}}{RT_2} + C + \frac{\Delta H^\circ_{\text{vap}}}{RT_1} - C$$

$$\ln P_2 - \ln P_1 = \frac{\Delta H^\circ_{\text{vap}}}{RT_1} - \frac{\Delta H^\circ_{\text{vap}}}{RT_2}$$

$$\ln \frac{P_2}{P_1} = \frac{\Delta H^\circ_{\text{vap}}}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

## 11.2 Vapor Pressure Clausius-Clapeyron Equation – Graphical Method

Determine enthalpy of vaporization graphically

**Question 1 of 2**

The vapor pressure of the liquid SO<sub>2</sub> is measured at different temperatures. The following vapor pressure data are obtained:

Temperature, K	Pressure, mmHg
220	81.55
230	147.39
240	253.55
250	417.68

Using the plotting tool, determine the enthalpy of vaporization,  $\Delta H_{\text{vap}}$ , and enter it in the box below.

kJ/mol

P  
 ln P  
 1/P

Least Squares Analysis

T  ln T  1/T

Determine enthalpy of vaporization graphically

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kJ/mol

P  
 ln P  
 1/P

Least Squares Analysis

Slope =  $-2.99 \times 10^3$

T  ln T  1/T

$$-\frac{\Delta H_{\text{vap}}}{R} = \text{Slope}$$

$$= -2.99 \times 10^3$$

$$\Delta H_{\text{vap}} = -2.99 \times 10^3 (8.314)$$

$$= -2.49 \times 10^4 \text{ J. mol}^{-1}$$

$$\Delta H_{\text{vap}} = 2.49 \times 10^4 \text{ J. mol}^{-1}$$

$$= 24.9 \text{ kJ. mol}^{-1}$$

## 11.2 Vapor Pressure

### Clausius-Clapeyron Equation – Quantative

From the following vapor pressure data for heptane, an estimate of the molar heat of vaporization of  $C_7H_{16}$  is

P, mm Hg	T, Kelvins
100	315
400	351

$$\ln \frac{P_2}{P_1} = \frac{\Delta H_{\text{vap}}^{\circ}}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$P_1 = 100 \quad T_1 = 315$$
$$P_2 = 400 \quad T_2 = 351$$

$$\ln \frac{400}{100} = \frac{\Delta H_{\text{vap}}^{\circ}}{R} \left( \frac{1}{315} - \frac{1}{351} \right)$$

$$\ln 4 (R) = \Delta H_{\text{vap}}^{\circ} (3.26 \times 10^{-4})$$

$$\Delta H_{\text{vap}}^{\circ} = \frac{1.39 (8.314)}{3.26 \times 10^{-4}}$$

$$= 3.54 \times 10^4 \text{ J. mol}^{-1}$$

OR

$$35.4 \text{ kJ. mol}^{-1}$$