Recitation 06

Mixture of Ideal Gases

1. Chapter 5: Exercise: 69

The partial pressure of CH_4 *(g)* is 0.175 atm and that of O_2 *(g)* is 0.250 atm in a mixture of the two gases.

a. **What is the mole fraction of each gas in the mixture?**
b. **If the mixture occupies a volume of 10.5 L at 65 ^{0}C, calculate the total number of moles of gas in the mixture.**
c. **Calculate the number of grams of each gas in the mixture.**

Solution:

(a) mol fraction of CH_4, $\chi_{CH4} = P_{CH4} / P_{total}$

$= 0.175 \text{ atm} / (0.175 \text{ atm} + 0.250 \text{ atm}) = 0.412$

$\chi_{O2} = 1.00 - \chi_{CH4} = 1.00 - 0.412 = 0.588$

(b) $PV = nRT$

$n_{total} = P_{total} V/RT$

$= (0.425 \text{ atm} * 10.5 \text{ L}) / (0.08206 \text{ L atm mol}^{-1}\text{K}^{-1} * 338 \text{ K})$

$= 0.161 \text{ mol}$

(c) $\chi_{CH4} = n_{CH4} / n_{total}$

$n_{CH4} = \chi_{CH4} * n_{total} = 0.412 * 0.161 \text{ mol} = 6.63 \times 10^{-2} \text{ mol } CH_4$

$= 6.63 \times 10^{-2} \text{ mol } CH_4 \times 16.04 \text{ g/mol } CH_4$

$= 1.06 \text{ g } CH_4$

$n_{O2} = 0.588 \times .161 \text{ mol} = 9.47 \times 10^{-2} \text{ mol } O_2$

$= 9.47 \times 10^{-2} \text{ mol } O_2 \times 32.0 \text{ g/mol } O_2$

$= 3.03 \text{ g } O_2$

Real Gases

2. Chapter 5: Exercise: 89

Calculate the pressure exerted by 0.5000 mol N_2 in a 1.000-L container at 25.0 ^{0}C

a. **Using the ideal gas law.**
b. **Using the van der Waals equation.**
c. **Compare the results**

Solution:

(a) Ideal gas law: $PV = nRT$
$P = nRT/V$
$= [0.500 \text{ mol} * 0.08206 \text{ L atm mol}^{-1}\text{K}^{-1} * (25 + 273.2)\text{K}] / 1.000 \text{ L}$
$= 12.24 \text{ atm}$

(b) van der Waals equation: $[P + a(n/V)^2] * (V – nb) = nRT$
For N_2, $a = 1.39 \text{ atm L}^2/\text{mol}^2$ and $b = 0.0391 \text{ L/mol}$
$[P + 1.39 (0.500/1.000)^2 \text{ atm}] * (1.000 \text{ L} – 0.5000*0.0391 \text{ L})$
$= 0.500 \text{ mol} * 0.08206 \text{ L atm mol}^{-1}\text{K}^{-1} * (25 + 273.2)\text{K}$

Or, $(P + 0.348 \text{ atm}) * (0.9805 \text{ L}) = 12.24 \text{ atm}$
$P = (12.24 \text{ L atm}/0.9805 \text{ L}) – 0.348 \text{ atm} = 12.48 – 0.348 = 12.13 \text{ atm}$

(c) The ideal gas law is high by 0.11 atm or (0.11/12.13)*100 = 0.91 %
Thus, using ideal gas law we got an over estimated value.

Intermolecular forces

3. Chapter 10: Exercise: 33

Rationalize the difference in boiling points for each of the following pairs of substances:

a. n-pentane, 36.2 ^{0}C
neopentane, 9.5 ^{0}C
b. HF 20 ^{0}C
HCl - 85 ^{0}C
c. HCl - 85 ^{0}C
LiCl 1360 ^{0}C
d. n-pentane 36.2 ^{0}C
n-hexane 69 ^{0}C

Solution:

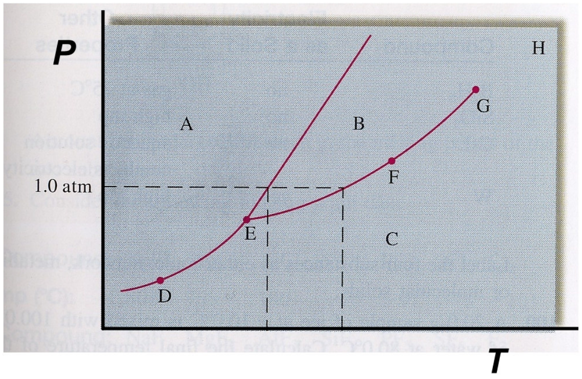
(a) Neopentane is more compact than n-pentane. Therefore, there is less surface area contact among neopentane molecules. This leads to weaker London dispersion forces and a lower boiling point.
(b) HF is capable of the hydrogen bonding, but HCl is not.
(c) LiCl is ionic and HCl is a molecular solid with only dipole forces and London dispersion forces. Ionic forces are much stronger than the forces for molecular solids.

(d) N-Hexane is a larger molecule, so it has stronger London dispersion force

Phase diagrams

4. Chapter 10: Exercise: 91

Consider the phase diagram given below. What phases are present at points A through H? Identify the triple point, normal boiling point, normal freezing point, and critical point. Which phase is denser, solid or liquid?



Solution:

A: Solid
B: Liquid
C: Gas/Vapour
D: Solid + Vapour
E. Solid + Liquid + Vapour
F: Liquid + Vapour
G: Liquid + Vapour
H: Vapur

The triple point is **E,** critical point is **G**
Normal freezing point: Temperature at which solid-liquid line is at 1 atm
Normal boiling point: Temperature at which liquid-vapour line is at 1 atm

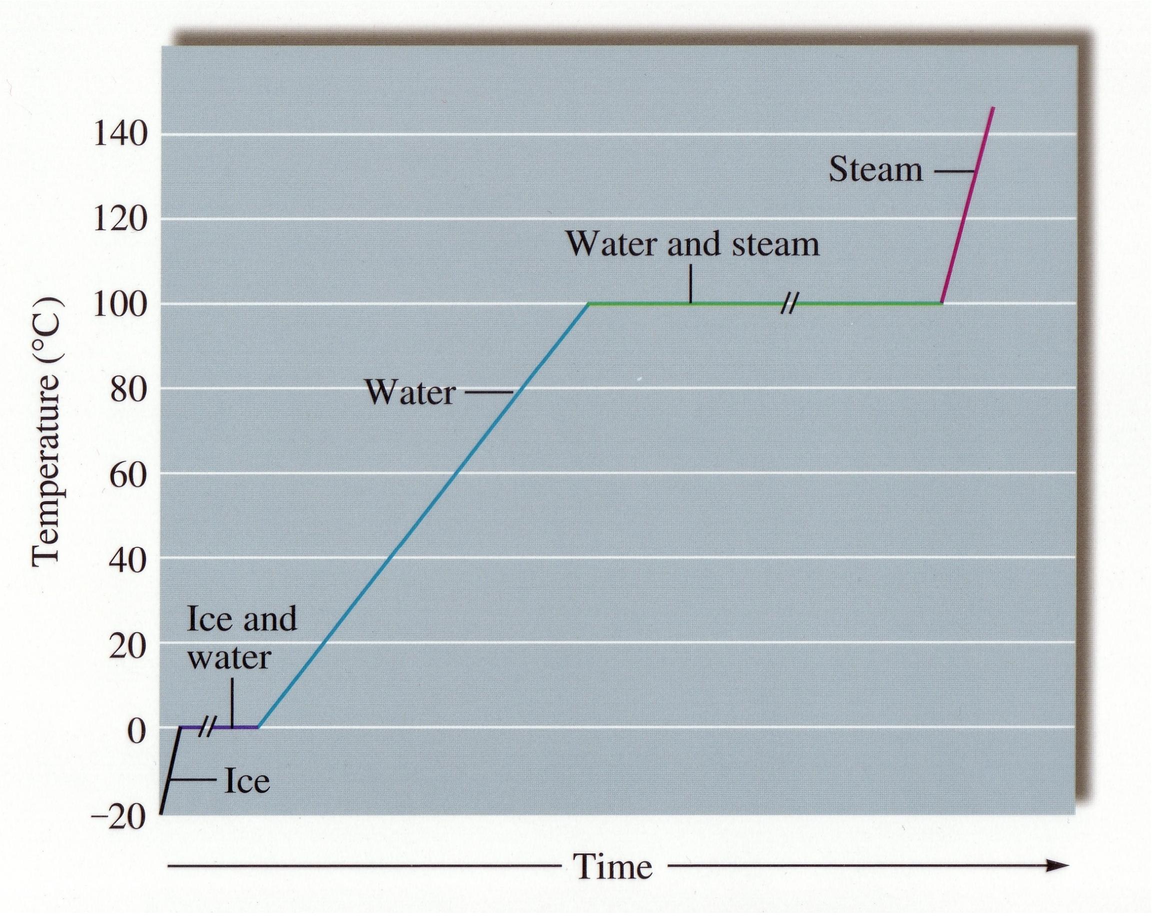
Since the solid-liquid line has a positive slope, the solid phase is denser than the liquid phase.

5. Heating curve of water is shown below:

When ice is heated continuously, the temperature increases with time. The heat evolved during this process is given as q (mass of the substance*heat capacity*change in temperature = m*C*ΔT).

The temperature of the system continues to increases until it reaches 0 ^{o}C when ice starts melting to form liquid water. During transformation of ice to liquid water at 0 ^{o}C, the temperature remains constant until all the ice is converted to liquid state. The heat evolved during the transition at 0 ^{o}C is called enthalpy of fusion ΔH_{fusion} (total heat evolved during the process is given as m*ΔH_{fusion}).

When all ice is converted to liquid water, again the temperature starts increasing till when liquid water starts converting to vapour at 100 ^{o}C. Similar to heat of fusion, the temperature remains constant at 100 ^{o}C until all the water in liquid state is converted to its vapour (steam). The heat evolved during this transition is called enthalpy of vaporization, ΔH_{fusion}. When all water gets converted to steam, again the temperature increases.



6. Application to heating curve:
Chapter 10: Exercise 87
How much energy does it take to convert 0.500 kh ice at – 20 ^{0}C to steam at 250 ^{0}C? Specific heat capacities: ice, 2.03 J/g.^{0}C; liquid, 4.2 J/g.^{0}C, steam, 2.0 J/g.^{0}C, ΔH_{vapour} = 40.7 kJ/mol, ΔH_{fus} = 6.02 kJ/mol.

Solution:

$$\underset{-20\,^{o}C}{H_2O\,(s)} \overset{q_1}{\text{-----------}} \underset{0\,^{o}C}{H_2O\,(s)} \overset{q_2}{\text{-----------}} \underset{0\,^{o}C}{H_2O\,(l)} \overset{q_3}{\text{-----------}} \underset{100\,^{o}C}{H_2O\,(l)} \overset{q_4}{\text{------------}} \underset{100\,^{o}C}{H_2O\,(g)} \overset{q_5}{\text{-----------}} \underset{250\,^{o}C}{H_2O\,(g)}$$

0.5 kg og ice = 500 g of ice = 500/18 mole of ice = 27.77 moles of ice

q_1 = m *C*ΔT = 500 g * 2.03 J/g. ^{o}C * 20 ^{o}C = 20 kJ
(ΔT = 0 ^{o}C – -20 ^{o}C = 20 ^{o}C)

$q_2 = m * \Delta H_{fusion} = 27.77 \text{ mol} * 6.02 \text{ kJ/mol} = 167 \text{ kJ}$

$q_3 = m * C * \Delta T = 500 \text{ g} * 4.2 \text{ J/g.}^{o}\text{C} * 100\ ^{o}\text{C} = 210 \text{ kJ}$
$(\Delta T = 100\ ^{o}\text{C} - 0\ ^{o}\text{C} = 100\ ^{o}\text{C})$

$q_4 = m * \Delta H_{vapour} = 27.77 \text{ mol} * 40.7 \text{ kJ/mol} = 1130 \text{ kJ}$

$q_5 = m * C * \Delta T = 500 \text{ g} * 2.0 \text{ J/g.}^{o}\text{C} * 150\ ^{o}\text{C} = 150 \text{ kJ}$
$(\Delta T = 250\ ^{o}\text{C} - 100\ ^{o}\text{C} = 150\ ^{o}\text{C})$

$q = q_1 + q_2 + q_3 + q_4 + q_5 = 20 \text{ kJ} + 167 \text{ kJ} + 210 \text{ kJ} + 1130 \text{ kJ} + 150\text{kJ}$
$= 1677 \text{ kJ}$