Chapter.1

THE PROPERTIES OF GASES

a) a) Could 131 g of xenon gas in a vessel of volume 1.0 dm^3 exert a pressure of 20 atm At 25°C if it behaved as a perfect gas? If not, what pressure would it exert?

b) What pressure would it exert if it behaved as a van der Waals gas?

(a) $P = \frac{nRT}{V}$

$= \frac{1mol \times 0.0821\ L.atm/mol.K\ \ 298k}{1\ L}$

$P = 24 atm$

$24 atm \neq 20\ atm$

(b) $(P + \frac{an^2}{V^2})(V - nb) = nRT$

$(P + \frac{4.137 \times 1}{1})(1 - 5.16 \times 10^{-2}) = 1 \times 0.0821 \times 298$

$P = 21.7 atm$

$\neq 20\ atm$

Pressure of real gas < pressure of ideal gas closed to the real pressure

2-a) A perfect gas undergoes isothermal compression, which reduced its volume by 2.20 dm^3. The final pressure and volume of the gas are 5.05 bar and 4.65 dm^3, respectively. Calculate the original pressure of the gas in (a) bar , (b) atm. 1bar = 0.987 atm

isothermal = const T

$\Delta V = V_2 - V_1$

$-2.2 = 4.65 - V_1$

$V_1 = 6.85\ dm^3$

$P_1 V_1 = P_2 V_2$

$P_1 \times 6.85 = 5.05 \times 4.65$

$P_1 = 3.4\ bar$

$P_1 = 3.4\ bar \times \frac{0.987\ atm}{1\ bar} = 3.355 atm$

8-a) At 500°C and 93.2 kPa, the mass density of sulfur vapour is 3.71 kg m^{-3}. What is the molecular formula of sulfur under these conditions?

$PV = \frac{m}{mwt} RT$

$Mwt = (\frac{m}{V}) \frac{RT}{P}$

$Mwt = \frac{3.71 \times 8.314 \times 773}{93.2 \times 10^3}$

$= 256 \times 10^{-3}\ Kg\ mol^{-1}$

$= 256\ g/mol$

$\frac{Mwt}{32} = 8$

S_8

$P_T = RH \times P_{given}$

9-a) Calculate the mass of water vapour present in a room of volume 400 m^3 that contains air at 27°C on a day when the relative humidity is 60 per cent.

1 atm = 1.013×10^5 Pa

1 atm = 760 torr

$P_{H_2O} = \frac{60}{100} \times 26.74$ torr = 16.04 Torr = 2137.9 Pa

$m = \frac{PV}{RT} Mwt$

$m = \frac{2137.9\ Pa \times 400\ m^3 \times 18 \times 10^{-3}\ kg\,mol^{-1}}{8.314\ J\,mol^{-1}\,K^{-1} \times 300\ K} = 6.2\ kg = 6173.1\ g$

13-a) Calculate the pressure exerted by 1.0 mol C_2H_6 behaving as H.W

(a) a perfect gas,

(b) a van der Waals gas when it is confined under the following conditions:

(i) at 273.15 K in 22.4 L; (ii) at 1000 K in 100 cm^3.

Use the data in Table 1.6.

(a) ① $P = \frac{nRT}{V} = \frac{1 \times 0.08206 \times 273.15}{22.4} = 1$ atm

② $P = \frac{nRT}{V} = \frac{1 \times 0.08206 \times 1000}{0.1} = 820.6$ atm

(b) ① $(P + \frac{an^2}{V^2})(V - nb) = nRT$

$(P + \frac{5.507 \times (1)^2}{(22.4)^2})(22.4 - 6.5 \times 10^{-2}) = (1)(0.08206)(273.15)$

$P = 0.993$ atm

② $(P + \frac{5.507 \times (1)^2}{(0.1)^2})(0.1 - 6.5 \times 10^{-2}) = (1)(0.08206 \times 1000$

$P = 1793.87$ atm

15-a) A gas at 250 K and 15 atm has a molar volume 12 per cent smaller than that calculated from the perfect gas law. Calculate:

(a) the compression factor under these conditions

(b) the molar volume of the gas.

Which are dominating in the sample, the attractive or the repulsive forces?

(b) $V_m = 1.20$

$Z = \frac{PV_m}{RT} = 0.88$

$Z < 1 \rightarrow$ attractive forces dominates

$Z > 1 \rightarrow$ repulsive forces dominates

16-a) In an industrial process, nitrogen is heated to 500 K at a constant volume of 1.000 m^3. The gas enters the container at 300 K and 100 atm. The mass of the gas is 92.4 kg. Use the van der Waals equation to determine the approximate pressure of the gas at its working temperature of 500 K.

For nitrogen, a = 1.352 dm^6 atm mol^{-2}; b = 0.0387 $dm^3 mol^{-1}$.

$n = \frac{m}{MWt} = \frac{92.4 \times 10^3}{28} = 3.30 \times 10^3$ mol

$P = \left(\frac{nRT}{V - nb}\right) - \left(\frac{an^2}{V^2}\right)$

$= \left(\frac{3.3 \times 10^3 \times 0.0821 \times 500}{1 \times 10^3 \text{ L} - 3.3 \times 10^3 \times 0.0387}\right) - \left(1.352 \times \frac{(3.3 \times 10^3)^2}{(1 \times 10^3)^2}\right)$

$= 140$ atm

18-a) A vessel of volume 22.4 L contains 2.0 mol H_2, and 1.0 mol N_2 at 273.15 K. Calculate

(a) the mole fractions of each component.

(b) their partial pressures, and

(c) their total pressure.

(a) $X_{H_2} = \frac{n_{H_2}}{n_T} = \frac{2}{3}$

$X_{N_2} = \frac{n_{N_2}}{n_T} = \frac{1}{3}$

(b) $P_{H_2} = \frac{nRT}{V} = \frac{2 \times 0.0821 \times 273.15}{22.4} = 2\,atm$

$P_{N_2} = \frac{nRT}{V} = \frac{1 \times 0.0821 \times 273.15}{22.4} = 1\,atm$

(c) $P_T = P_{H_2} + P_{N_2}$

$= 2 + 1 = 3\,atm$

19 -a) The critical constants of methane are p_c = 45.6 atm. V_c = 98.7 $cm^3\ mol^{-1}$ and T_c = 190.6 K. Calculate the van der Waals parameters of the gas and estimate the radius of the molecules.

$V_m = b = 32.9\ cm^3/mol$

$V_{molec} = \frac{b}{N_A} = 5.46 \times 10^{-23}\ cm^3$

$V_{molecule} = \frac{4}{3}\pi r^3$

$r = 2.4 \times 10^{8}\ cm$

$= 0.24\ nm$

20-a) Use the van der Waals parameters for chlorine to calculate approximate values of

(a) the Boyle temperature of chlorine and

(b) the radius of a Cl_2 molecule regarded as a sphere.

21-a) Suggest the pressure and temperature at which 1.0 mol of a) NH_3, b) Xe, c) He will be in states that correspond to 1.0 mol H_2 at 1.0 atm and 25°C.

22-a) A certain gas obeys the van der Waals equation with b =0.50 m^6 Pa mol^{-2}. Its volume is found to be 5.00 x 10^{-4} m^3 mol^{-1} at 273 K and 3.0 MPa. From this information calculate the van der Waals constant b. What is the compression factor for this gas at the prevailing temperature and pressure?

$$\left(P + \frac{an^2}{V^2}\right)(V - bn) = nRT$$

$$b = V - \frac{R\ T}{P + \frac{an^2}{V^2}}$$

$$= 5 \times 10^{-4} - \frac{8.314 \times 273}{3 \times 10^6 + \left(\frac{0.5}{(5 \times 10^{-4})^2}\right)}$$

$$b = 4.6 \times 10^{-5}\ m^3/mol$$

1-a) Determine the ratios of

(a) the mean speeds

(b) the mean kinetic energies of H_2 molecules and Hg atoms at $20°c$

$$\rightarrow C = \left(\frac{8RT}{\pi M}\right)^{1/2}$$

$$= \left(\frac{8 \times 8.314 \times 293}{\pi \times (200.6 \times 10^{-3})}\right)^{1/2}$$