The Copperbelt University

CH 110 Tutorial Sheet 4 – October 2, 2015

Gases

Question 1

Good understanding of gas concepts requires that basic information such as *pressure units*, gas laws, *stoichiometry* and mole fraction are properly understood.

a) Define the terms in *italics*.

ANSWER:

Pressure Units: ***These are derived units of parameter Pressure. Pressure is defined as Force per unit. SI units, Pressure is measured in Newtons per square metre.***

Stoichiometry: ***Thisis the quantitative study of reactants and products in a chemical reaction.***

b) Give four examples of pressure units and show how they are related to each other.

ANSWER:

(i) Newtons per square metre (1 N/m^2=1 Pa)

(ii) Pascal

(iii) Torr (1 mm Hg = 1 Torr)

(iv) Atmospheres (1 atm = 760 Torr = 101325 Pa)

(v) Bars (1 Bar = 100,000 Pa)

(vi) mm Hg (1 atm=760 mm Hg)

c) What is mole fraction? How is it calculated mathematically?

ANSWER:

Mole fraction the ratio of the number of moles of a given component in a mixture to the total number of moles in the mixture.

Mathematically it is calculated as $x_i = \frac{n_i}{n_T}$

Question 2

Convert the following volumes of gas into volumes at s.t.p:

a) 30 cm^3at 261 K and 77 kPa

ANSWER:

Using the Gas Law $\frac{P_1V_1}{T_1} = n \times R = \frac{P_2V_2}{T_2}$ ***as Equation 1 where V_1=30 cm^3, P_1=77 kPa, T_1=261K, P_2=1 atm= 101.325 kPa and T_2=273 K***

Making V_2 the subject gives equation 2 which is $V_2 = \left(\frac{P_1 \times V_1}{T_1}\right) \times \left(\frac{T_2}{P_2}\right)$.

Substituting the given values gives

$$V_2 = \left(\frac{77\ kPa \times 30\ cm^3}{261\ K}\right) \times \left(\frac{273\ K}{101.325\ kPa}\right) = 8.85 \times 2.69 = \underline{\underline{23.8\ cm^3}}$$

b) 450 cm^3at 303 K and 102.7 KPa

ANSWER:

Using the Gas Law $\frac{P_1V_1}{T_1} = n \times R = \frac{P_2V_2}{T_2}$ ***as Equation 1 where V_1=450 cm^3, P_1=102.7kPa, T_1=303 K, P_2=1 atm= 101.325 kPa and T_2=273 K***

Making V_2 the subject gives equation 2 which is $V_2 = \left(\frac{P_1 \times V_1}{T_1}\right) \times \left(\frac{T_2}{P_2}\right)$.

Substituting the given values gives

$$V_2 = \left(\frac{102.7\ kPa \times 450\ cm^3}{303\ K}\right) \times \left(\frac{273\ K}{101.325\ kPa}\right) = 152.52 \times 2.69 = \underline{\underline{412\ cm^3}}$$

c) 50 cm^3at 620 K and 24.312 x 10^5Pa

ANSWER:

Using the Gas Law $\frac{P_1V_1}{T_1} = n \times R = \frac{P_2V_2}{T_2}$ ***as Equation 1 where V_1=50 cm^3, P_1=2431.2kPa, T_1=620 K, P_2=1 atm= 101.325 kPa and T_2=273 K***

Making V_2 the subject gives equation 2 which is $V_2 = \left(\frac{P_1 \times V_1}{T_1}\right) \times \left(\frac{T_2}{P_2}\right)$.

Substituting the given values gives

$$V_2 = \left(\frac{2431.2\ kPa \times 50\ cm^3}{620\ K}\right) \times \left(\frac{273\ K}{101.325\ kPa}\right) = 196.06 \times 2.69 = \underline{\underline{528\ cm^3}}$$

d) Determine the volume of O_2 (at s.t.p) formed when 50.0 g of $KClO_3$ decomposes according to the following reaction:

$$2\ KClO_3(s) \longrightarrow 2\ KCl\ (s) + 3\ O_2(g)$$

ANSWER:

When 2 moles of $KClO_3$ decomposes STP gives 3 moles of O_2 or 1 mole of $KClO_3$ decomposes STP gives 1.5 moles of O_2

The molar mass of $KClO_3$ is

$$M_r = M_K + M_{Cl} + 3 \times M_O = 39.10 + 35.35 + 48.00 = 122.55\ g$$

Thus, the number of moles of $KClO_3$ in 50 g is $n = \frac{m_s}{M_r} = \frac{50}{122.55} = 0.408\ mol$

The number of moles of oxygen formed at its decomposition is

0.408×1.5 mole = $\underline{\underline{0.612\ moles}}$

Since 1 mole of gas at STP occupies 22.4 L, 0.612 moles will occupy

0.612×22.4 L = $\underline{\underline{13.7\ L}}$

Question 3

a) A sample of gas originally occupies 29.1 L at 0.0 °C. What is the new volume when it is heated to 15°C? (Assume the pressure is constant)

ANSWER:

Using the Gas Law $\frac{P_1V_1}{T_1} = n \times R = \frac{P_2V_2}{T_2}$ ***as Equation 1 where*** V_1***=29.1 L,*** P_1***=P kPa,*** T_1***=273 K,*** P_2***=P kPa and*** T_2***=288 K***

Making V_2 ***the subject gives equation 2 which is*** $V_2 = \left(\frac{P_1 \times V_1}{T_1}\right) \times \left(\frac{T_2}{P_2}\right)$***. Since*** P_1***=***P_2***=P kPa, the above equation simplifies to*** $V_2 = \frac{V_1 \times T_2}{T_1}$

Substituting the given values gives

$$V_2 = \frac{29.1 \times 288}{273} = \frac{8380.9}{273} = \underline{\underline{30.7L}}$$

b) Given pressure in units of Pascal, volume in units of cubic metres, temperature in units of Kelvin and amount of substance in units of moles; derive the units of R, the universal gas constant from the ideal gas equation.

ANSWER:

Using the Ideal Gas Law expression we have $PV = nRT$ ***as Equation 1 where V is in m***^{3}***, P is in Pa, T is K, n is in mol***

Making R the subject gives equation 2 which is $R = \frac{P \times V}{n \times T}$***.***

Since Pascal (Pa) is as derived unit is defined of force per unit area, its units are Newtons per metre squared (m^2***).***

The Newton is derived of mass by acceleration, whose derived units are

$$kg \times m \times s^{-2}$$

The unit of Pressure which is force per unit area becomes

$$\frac{kg \times m \times s^{-2}}{m^2} = kg \times m^{-1} \times s^{-2}$$

The product PV has the units

$$\left(kg \times m^{-1} \times s^{-2}\right) \times m^3 = kg \times m^2 \times s^{-2}$$

This is the same as the units of Force by Distance which is the derived unit of Work which is called the Joule (J).

The product nT has the units mol K.

Therefore the unit of R is J/(mol K) or $\underline{\underline{Jmol^{-1}K^{-1}}}$

c) A series of measurements were made in order to determine the molar mass of anunknown gas, **M**. Firstly, a flask was evacuated and found to weigh 134.567 g. It wasthen filled with the gas to a pressure of 735 torr at 31°C and reweighed; its masswas then 137.456 g. Finally, the flask was filled with water at 31°C and found toweigh 1067.9 g. Assuming that the ideal gas equation applies, calculate the molarmass of the unknown gas, **M**. (The density of water at 31°C was 0.997 g/mL)

ANSWER

Using the Ideal Gas Law expression we have $PV = nRT$ ***as Equation 1***

Since $n = \frac{m_s}{M_r}$***, when this is placed in the ideal gas we obtain equation 2***

$PV = \frac{m_s}{M_r}RT$. Making molar mass the subject of the formula yields equation 3

$M_r = \frac{m_s RT}{PV}$. When volume of the container is indirectly determined using mass and density of water we can replace V by equation 4 which is $V = \frac{m_{H_2O}}{\rho_{H_2O}}$. Substituting 4 into 3 gives equation 5

$$M_r = \frac{m_s RT}{P\times\frac{m_{H_2O}}{\rho_{H_2O}}} = \frac{\rho_{H_2O}\times m_s\times R\times T}{P\times m_{H_2O}}.(5)$$

where

$m_s = \Delta m = mass\ of\ gas\ filled\ flask - mass\ of\ empty\ flask or$

$\Delta m = 137.456\ \text{g} - 134.567\ \text{g} = 2.889\ \text{g},$

density of water $\rho_{H_2O} = 0.997\frac{g}{mL} = 977\ kg/m^3$;

T = 31$°C = 304\ K$; ***P= 735 Torr*** $=\frac{735}{760}\times 101325\ Pa = 97992\ Pa$ ***(or Nm^{-2})***

m_{H_2O}***=mass of full flask – empty flask =***$1067.9\ \text{g} - 134.567\ \text{g} = 933.333\ \text{g}$

$R = 8.314\ Jmol^{-1}K^{-1}$

Substituting the relevant values gives

$$M_r = \frac{997kg\ m^{-3}\times 2.889g\times 8.314\ J\ mol^{-1}K^{-1}\times 304\ K}{97992\ N\ m^{-2}\times 0.933333\ kg} = 79.6\ g/mol$$

Question 4:

a) Explain the term ‘partial pressure’.

ANSWER:

Partial pressures of a gas mixture are the independent pressures exerted by different gases of the mixture. For example if we have a mixture of three gases, $P_{Total}=(P_1+P_2+P_3)=(n_1+n_2+n_3)RT$ where P_1, P_2 and P_3 are the partial pressures of three gases.

We can also express partial pressures as $P_{Total}=(\chi_1+\chi_2+\chi_2)P_{Total}$ where $\chi_i = \frac{P_i}{P_{Total}} = \frac{n_i}{n_{Total}}$ is the mole fraction

b) The mass percentage composition of dry air at sea level is approximately: N_2: 75.5; O_2: 23.2; Ar: 1.3. Calculate the partial pressure of each component when the total pressure is 1.5 atm.

ANSWER:

We need to calculate the mole fraction each gas and use it to calculate the partial pressure of each gas.

We give P_{Total} = 1.5 atm

If we assume sample is 100 g,

Gas	*Molar Mass (M_r)*	*Mass (m_s)*	*Moles (n_i)*
Nitrogen	*28.02 g*	*75.5 g*	$n_1 = \frac{75.5}{28.02} = 2.6945$
Oxygen	*32.00 g*	*23.2 g*	$n_2 = \frac{23.2}{32.00} = 0.7250$
Argon	*39.95 g*	*1.3 g*	$n_3 = \frac{1.3}{39.95} = 0.0325$
$n_{Total} = n_1 + n_2 + n_3$			*3.452*

The values of mole fraction and partial pressures are given in the table below

Gas	*Mole fraction* ($\chi_i = \frac{n_i}{n_1+n_2+n_3}$)	*Partial Pressure is* $P_i = \chi_i \times P_{Total}$
Nitrogen	$\chi_1 = \frac{2.6945}{3.452} = 0.78$	*0.78×1.5 =1.171 atm*
Oxygen	$\chi_2 = \frac{0.725}{3.452} = 0.21$	*0.21×1.5=0.315atm*
Argon	$\chi_3 = \frac{0.0325}{3.452} = 0.01$	*0.01×1.5=0.015 atm*

c) A vessel of volume 5 $dm^3$was filled with:

i. 2 mol of H2 (g)
ii. Then a further 1 mol of Cl2 (g)
iii. The mixture was sparked leading to the formation of 2 mol of HCl (g)
What is the partial pressure of the component gases at each stage and the total pressure, the temperature being 25 °C throughout?

ANSWER:

The chemical reaction involved is

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$$

The amounts of reagents at different stages of the reaction are shown in the table below

	$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$
t=0	2 mol		1 mol		0 mol
t=end	1 mol		0 mol		2 mol

Partial pressure is given by the equation $P_i = \frac{n_i \times RT}{V}$ ***while*** $P_{Total} = \frac{(n_1+n_2+n_3)\times RT}{V}$

With R=8.314 J/(mol K), T=298 K and V=0.005 m³, then the value of $\frac{RT}{V}$ is given as

$$\frac{RT}{V} = \frac{8.314 \times 298}{0.005} = 495.5\ kPa/mol$$

So that evaluated pressures are shown in the table below:

Stage	*Gas*	*n*	*Partial or Total Pressure*
t=0	H_2	*2*	$P_{H_2} = \frac{n_{H_2}RT}{V} = 2 \times 495.5 = 991\ kPa$
	Cl_2	*1*	$P_{Cl_2} = \frac{n_{Cl_2}RT}{V} = 1 \times 495.5 = 495.5\ kPa$
	HCl	*0*	*0*
	Total	*3*	$P_{Cl_2+H_2} = \frac{(n_{Cl_2} + n_{H_2})RT}{V} = 3 \times 495.5 = 1486.5\ kPa$
t=end	H_2	*1*	$P_{H_2} = \frac{n_{H_2}RT}{V} = 1 \times 495.5 = 495.5\ kPa$
	Cl_2	*0*	*0*
	HCl	*2*	$P_{HCl} = \frac{n_{HCl}RT}{V} = 2 \times 495.5 = 991\ kPa$
	Total	*3*	$P_{H_2+HCl} = \frac{(n_{HCl} + n_{H_2})RT}{V} = 3 \times 495.5 = 1486.5\ kPa$

Question 5:

a) Calculate the rate of effusion of methane to that of sulphur dioxide.

ANSWER:

According to Graham's law of effusion,

$$\frac{Rate\ of\ effusion\ of\ Gas\ 1}{Rate\ of\ effusion\ of\ Gas\ 2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

where M_1 and M_2 represent the molar masses of the gases.
In the case where M1 is the molar of methane =16.04 g/mol and M2 is the molar mass of sulphur dioxide = 64.00g/mol, then

$$\frac{Rate\ of\ effusion\ of\ methane}{Rate\ of\ effusion\ of\ sulphur\ dioxide} = \frac{\sqrt{M_{SO_2}}}{\sqrt{M_{CH_4}}} = \frac{\sqrt{64.00}}{\sqrt{16.04}} \approx \frac{8}{4} = 2$$

b) A certain volume of the gas evolve by photosynthesizing alga culture took 231 s to stream through a small hole. Under precisely the same conditions, an equal volume of argon took 258 s. Calculate the molar mass of the unknown gas.

ANSWER:

According to Graham's Law of effusion, we have

$$\frac{t_x}{t_{Argon}} = \frac{\sqrt{M_{Argon}}}{\sqrt{M_x}}$$

Squaring both sides and making M_x the subject, we have

$$M_x = M_{Argon} \times \left(\frac{t_{Argon}}{t_x}\right)^2$$

Substituting the M_{Argon}= 39.95 g/mol, t_x=231 s and t_{Argon}=258 s, we have

$$M_x = 39.95 \times \left(\frac{258}{231}\right)^2 = 49.8\ g/mol$$

c) Gas **X** effuses at the rate that is 0.355 times the rate of O_2under the same conditions. Propose the name of gas **X**.

ANSWER:

According to Graham's Law of effusion, we have

$$\frac{t_x}{t_{Oxygen}} = \frac{\sqrt{M_{Oxygen}}}{\sqrt{M_x}}$$

Given that t_x/t_{Oxygen} = 0.355 and M_{Oxygen} = 32 g

Substituting these values in the equation and solving for M_x we have

$$M_x = 32 \times \left(\frac{1}{0.355}\right)^2 = 253.9\ g/mol$$

There is no monoatomic gas with this molar mass but a diatomic gas with this molar mass is Iodine Vapour.

Question 6

a) Explain the postulates of the kinetic molecular theory of gases.
ANSWER:
Postulates of the KMT:
- ***(i) Volume of gas particles is zero***
- ***(ii) No particle interactions***
- ***(iii) Particles are in constant motion, colliding with the container walls to produce pressure***
- ***(iv) The average kinetic energy of the gas particles is directly proportional to the temperature of the gas in kelvins***

b) Outline two assumptions made about ideal gases that do not apply to real gases which result in real gases deviating from the ideal gas behaviour.
ANSWER:
Assumptions made about ideal gases that do not apply to real gases are the same as the first two postulates of the KMT, namely:
- ***(i) Volume of gas particles is zero***
- ***(ii) No particle interactions***

c) Show how the ideal gas equation can be modified to account for the behaviour of real gases.
ANSWER:
Required corrections are that real gases
- ***(i) have volumes proportional to the number of moles so that V'=V-nb, where V' is volume of the container (V) less the volume (nb)of the gas where n is the number of moles of the gas and b is an empirical constant***
- ***(ii) have a pressure interaction term that is proportional to square of the number of moles per unit volume. The proportionality constant (a) is also an empirical factor.***

The corrected ideal gas law equation is the van der Waals equation and is given as

$$\underset{\text{Observed Pressure}}{P_{obs}} = \frac{nRT}{\underset{\text{Volume of the container}}{V} - \underbrace{nb}_{\text{Volume correction}}} - \underbrace{a\left(\frac{n}{V}\right)^2}_{\text{Pressure correction}}$$