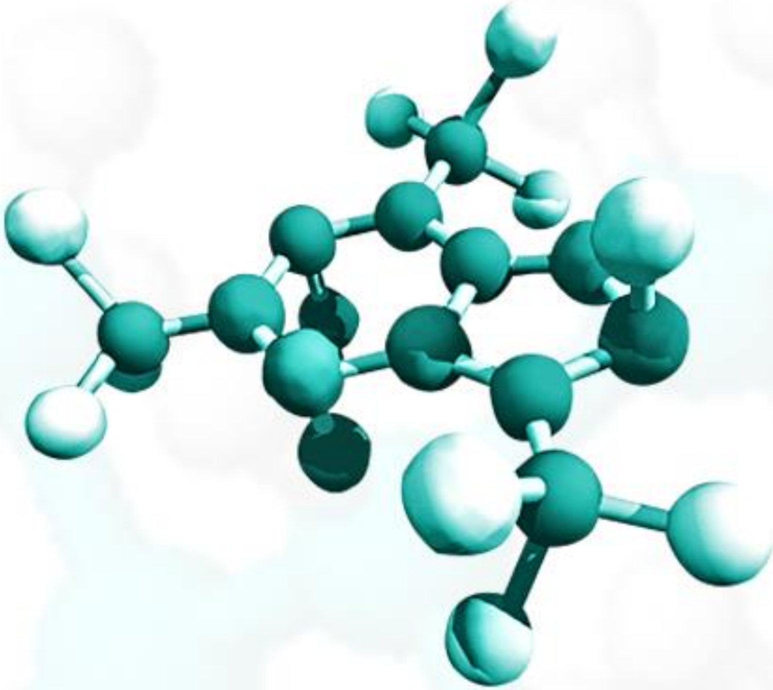
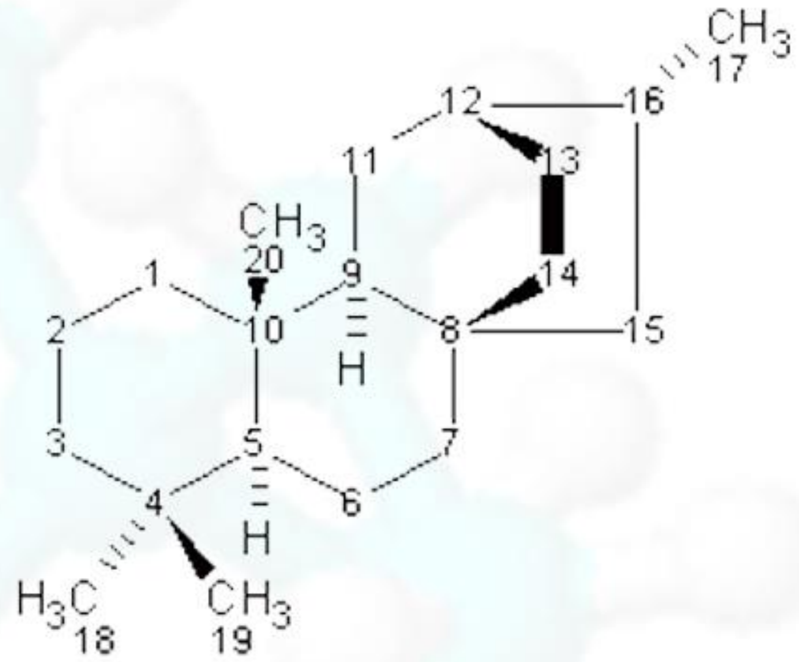
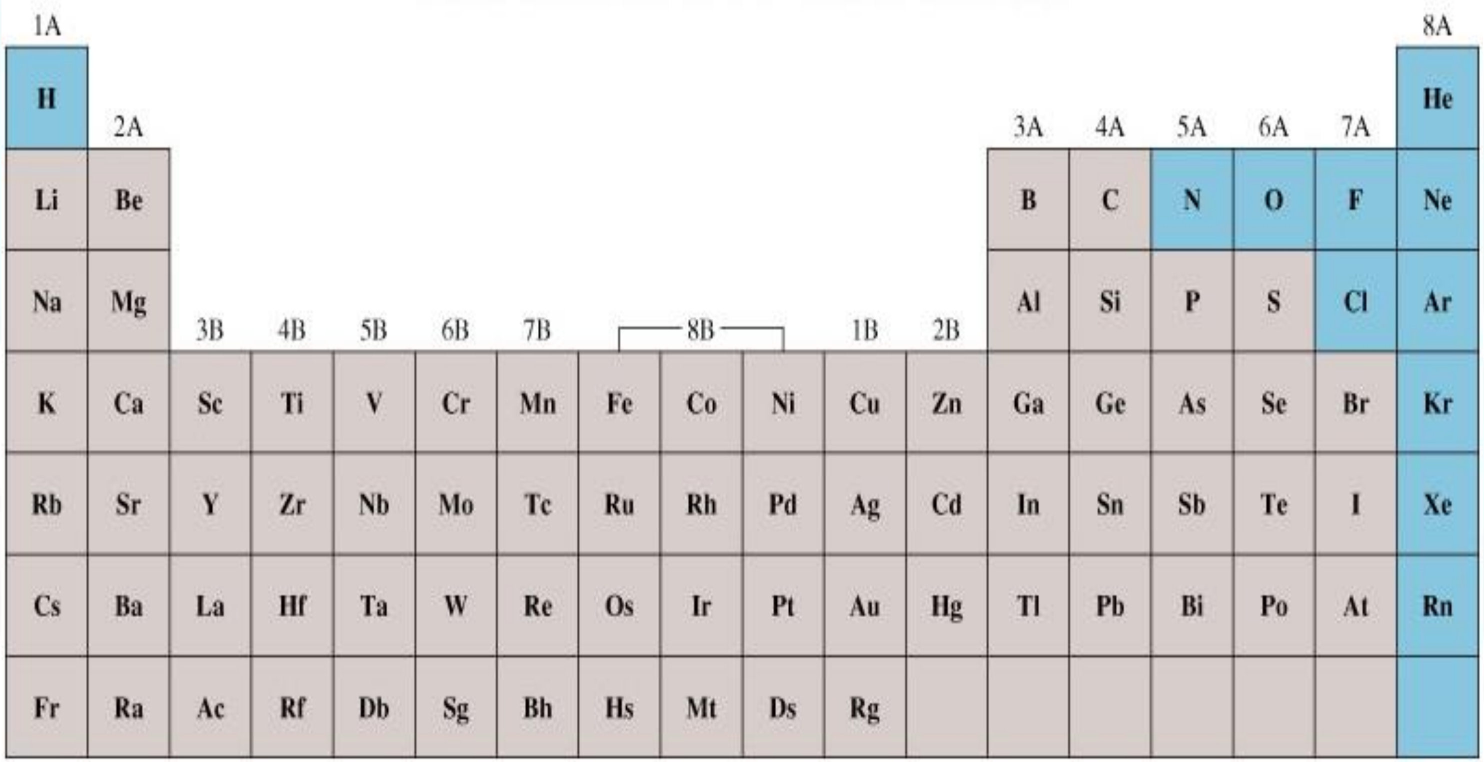




Chapter Five

Gases

Substances That Exist as Gases



- Element in blue are Gases
- Noble gases are monatomic
- All other gases (H_2, N_2, O_2, F_2, Cl_2) diatomic molecules.

Substances That Exist as Gases

TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

Gas pressure

Unites of pressure :

Pascal (Pa), atm, mmHg, torr

1 torr = 1mmHg

1 atm = 760 mmHg

1 atm = 1.01325×10^5 Pa.

Example1 :

convert the pressure of 688 mmHg to atmospheric pressure?

1 atm = 760mmHg

X

? atm = 688 mmHg

760 X ? = 1 X 688

Pressure = 688 / 760 = 0.905 atm.

Gas Laws

- For every gas there are :

P (pressure), T (Temperature) V (volume), n (mole number).

The Pressure – Volume Relationships
Boyle's Law

- Boyles's law study the relationship between the pressure and volume of gas.
- Boyel's law stated that *the pressure of a fixed amount of gas at a constant temperature in inversely proportional to the volume of the gas.*

$$P \,\alpha\, \frac{1}{V}$$

Gas Laws- Boyle's Law

$$P_1 \; \alpha \; \frac{1}{V_1} \Rightarrow P_1 = \frac{k}{V_1} \Rightarrow k = P_1 \; x \; V_1$$

$$P_2 \; \alpha \; \frac{1}{V_2} \Rightarrow P_2 = \frac{k}{V_2} \Rightarrow k = P_2 \; x \; V_2$$

$P_1 V_1 = P_2 V_2$ ← Boyle's Law

Gas Laws- Boyle's Law

Example 1 :

A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL?

P_1 = 726 mmHg, V_1= 946 ml, P_2= ?, V_2= 154 mL.

$P_1 V_1 = P_2 V_2$

726 x 946 = P_2 x 154

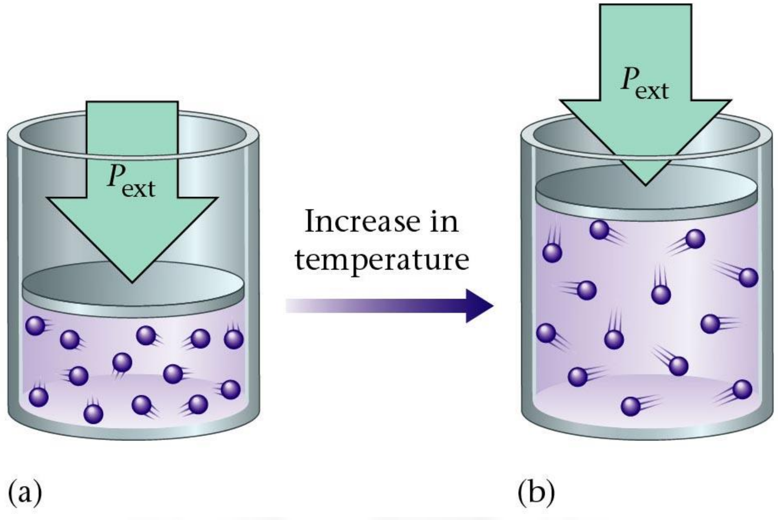
$$P_2 = \frac{726 \text{ x } 946}{154}$$

$$= 4459.7 \text{ mmHg}$$

Gas Laws

The Temperature – Volume Relationships Charle's and Gay-Lussac's Law

- Charle's and Gay-Lussac's law study the relationship between the temperature and volume of gas.
- Charle's and Gay-Lussac's law stated that the volume of a fixed amount of gas at a constant pressure is directly proportional to the *absolute temperature* of the gas.



Gas Laws- Charle's Law

T α V

$$T_1 \ \alpha \ V_1 \quad \Rightarrow T_1 = k \times V_1 \Rightarrow k = \frac{T_1}{V_1}$$

$$T_2 \ \alpha \ V_2 \quad \Rightarrow T_2 = k \times V_2 \Rightarrow k = \frac{T_2}{V_2}$$

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

Charle's Law

T in Kelvin

Gas Laws- Charle's Law

Example:

A sample of carbon monoxide gas occupies 3.20 L at 125 °C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

$$\frac{125 + 273}{3.2} = \frac{T_2}{1.54}$$

T_2 X 3.2 = 398 X 1.54

T_2 = 612.92/3.2

= 191.5 K

Gas Laws

The Volume – Amount Relationships
Avogadro's Law

- Avogadro's law study the relationship between the volume and number of mole of gas.
- Avogadro's law stated that at constant pressure and temperature, the volume is directly proportional to the number of moles of the gas

Gas Laws- Avogadro's Law

$n \propto V$

$$n_1 \propto V_1 \Rightarrow n_1 = k \times V_1 \Rightarrow k = \frac{n_1}{V_1}$$

$$n_2 \propto V_2 \Rightarrow n_2 = k \times V_2 \Rightarrow k = \frac{n_2}{V_2}$$

$$\frac{n_1}{V_1} = \frac{n_2}{V_2}$$

← Avogadro's Law

Summery of Gas Laws

$P_1 V_1 = P_2 V_2$

Boyle's Law
Constant T and n

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

Charle's Law
Constant P and n

$$\frac{n_1}{V_1} = \frac{n_2}{V_2}$$

Avogadro's Law
Constant P and T

Ideal Gas Equation

We know that

$V \,\alpha\, \frac{1}{P}$ Boyle's law

$V \,\alpha\, T$ Charle's law

$V \,\alpha\, n$ Avogadro law

Then

$$V \,\alpha\, \frac{nT}{P}$$

$$V = \frac{nRT}{P}$$

Ideal Gas Equation

PV=nRT

P= pressure (atm), V= volume (L), n= moles
R= gas constant, T= temperature (K)

Ideal Gas Equation

- Ideal gas is a hypothetical gas whose pressure-volume-temperature behavior can be completely accounted for by the ideal gas equation.
- STP : standard Temperature and pressure
- Standard Temperature = 0°C = 273.15 K
- Standard Pressure = 1 atm.
- At STP 1mole of an ideal gas occupies 22.414L.
- R (gas constant) = 0.0821 L.atm / K.mol

Ideal Gas Equation

Example 1:

Calculate the pressure (in atm) exerted by 1.82 moles of the sulphur hexaflouride in a steel vessel of volume 5.43 L at 69.5 °C.?

PV =nRT

P = nRT/V

= 1.82 x 0.0821 x (69.5 +273)/5.43

=9.41 atm.

Ideal Gas Equation

Example 2:

Calculate the volume (in liters) occupied by 7.40g of NH_3 at STP condition.?

PV=nRT

n = mass/molar mass

n = 7.40 / 17 = 0.435 mol

V = nRT/P

V= 0.435 X 0.082 X 273 / 1

= 9.74 L

Ideal Gas Equation

- We can use the ideal gas law if we know three out of four variable namely: P,T,V,n. we can calculate one unknown if we know the other three from the equation of ideal gas.
- However, sometime we have to deal with two conditions, this means we have two P , two V, two T, and two n. thus we need to apply some modification into the equation of ideal gas that take into account the initial and final conditions.

PV = nRT $$R = \frac{PV}{nT}$$

$$R = \frac{P_1 V_1}{n_1 T_1} \text{ befor change}$$

$$R = \frac{P_2 V_2}{n_2 T_2} \text{ after change}$$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

Normally $n_1=n_2$

And the law become

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Ideal Gas Equation

Example 1:

A small bubble rises from the bottom of a lake, where the temperature and pressure are 8 °C and 6.4 atm, to the water surface, where the temperature 25 °C and the pressure is 1 atm. Calculate the final volume (in mL) of the bubble if its initial volume was 2.1 mL?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

$$V_2 = \frac{6.4 \text{ x } 2.1 \text{ x } (25+273)}{1 \text{ x } (8+273)}$$

V_2 = 14.25 mL

Ideal Gas Equation

Example 2:

An inflated helium balloon with a volume of 0.55L at sea level (1 atm) is allowed to rise to a high of 6.5 km. where the pressure is about 0.40 atm. Assuming that the temperature remains constant. What is the final volume of the balloon?

We assume that $n_1 = n_2$ and $T_1 = T_2$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

$$P_1 V_1 = P_2 V_2$$

$$1 \times 0.55 = 0.4 \times V_2$$

$$V_2 = 0.55 / 0.4$$

$$= 1.4L$$

Density Calculations

P V = n R T

$$P = \frac{n}{V} R T$$

$$\frac{n}{V} = \frac{P}{R T}$$

I know that n = mass /molar mass

$$\frac{m}{MM\,V} = \frac{P}{R T}$$

I know d= mass/volume

$$\frac{d}{MM} = \frac{P}{R T}$$

P MM = d R T

$$d = \frac{P\,MM}{R T}$$

Unite for gas density is g/L

d = density (g/L), P= pressure (atm) MM= molar mass (g/mol), R= gas constant, T= temperature (K)

Density Calculations

Example 1:

Calculate the density of CO_2 in g/L at 0.990 atm and 55 °C?

$$d = \frac{P\,MM}{R\,T}$$

$MM(CO_2)$= 40 g/mol

$$d = \frac{0.99 \text{ x } 40}{0.0821 \text{ x } (55 + 273)}$$

d= 1.47 g/L

The molar mass of a gaseous substance

- Normally we can determine the molar mass of a compound from the chemical formula.
- However, sometime we work with unknown compound or partially known compound. If the unknown substance is gaseous, its molar mass can be determine from the ideal gas equation. All needed is the density of the gas (or mass and volume of the gas).

$$d = \frac{P\,MM}{R\,T}$$

$$d\,R\,T = P\,MM$$

$$MM = \frac{d\,R\,T}{P}$$

The molar mass of a gaseous substance

Example 1:

A chemist has synthesised a green-yellow gaseous compound of chlorine and oxygen and finds that its density is 7.71 g/L at 36 °C and 2.88 atm. Calculate the molar mass of the compound?

$$MM = \frac{d\,R\,T}{P}$$

$$MM = \frac{7.71 \times 0.0821 \times (36 + 273)}{2.88}$$

$$MM = 67.9\ g/mol$$

The molar mass of a gaseous substance

Example 2:

Chemical analysis of a gaseous compound showed that it contained 33.0 percent Si and 67.0 percent F by mass. At 35 °C, 0.210 L of the compound exerted a pressure of 1.70 atm. If the mass of 0.210 L of the compound was 2.38 g, calculate the molar mass and determine the molecular formula of the compound?

Si = 33%, F= 67%, T= 35 °C, V= 0.210L, P= 1.7 atm, mass= 2.38g,

MM= ?, Molecular formula ??

$$MM = \frac{dRT}{P}$$

$$d = \frac{m}{V}$$

$$d = \frac{2.38}{0.210}$$

d=11.33 g/L

$$MM = \frac{11.33 \times 0.0821 \times (35 + 273)}{1.70}$$

MM=168.5 g/mol

The molar mass of a gaseous substance

1- change from % to g

33 g of Si, 67 g of F,

2- change from g to mole using

$$n_{Si} = \frac{33}{28.09} = 1.17 \text{ mol of Si}$$

$$n_F = \frac{67}{19} = 3.53 \text{ mol of F}$$

Divided by the smallest number of mole which is 1.17

$$\text{Si: } \frac{1.17}{1.17} = 1 \qquad \text{F: } \frac{3.53}{1.17} \approx 3$$

Thus the empirical formula is SiF_3

Then we calculate the molar mass of the empirical formula

SiF_3= 85.09 g/mol

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

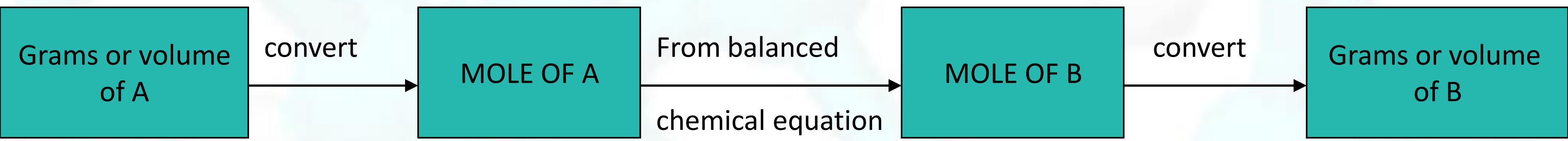
$$\text{Ratio} = \frac{168.5}{85.09} \approx 2$$

Molecular formula = empirical formula x ratio

= SiF_3 x 2 = Si_2F_6

Gas Stoichiometry

- In chapter 3 we learned how to calculate the product amount if we know the amount of reactant or how to calculate the amount of reactant if know the amount of product.
- The relationship was between n and m.
- In gases we can do the same however the relationship is between V and n.



We can use the volume only when the product and reactant are gases. And when T and P are constant

Gas Stoichiometry

Example 1:

Calculate the volume of O_2 (in L) required for the complete combustion of 7.64 L of C_2H_2 measured at the same temperature and pressure.?

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

From equation

2 mol C_2H_2 $\longrightarrow$ 5 mol O_2

2L C_2H_2 $\longrightarrow$ 5L O_2

7.64 L C_2H_2 $\xrightarrow{X}$? L O_2

5 x 7.64 = 2 x ?

Volume of O_2 = 5 X 7.64 / 2 = 19.1 L

Gas Stoichiometry

Example 2:

Sodium azide (NaN_3) is used in some automobile air bags. The impact of a collision triggers the decomposition of NaN_3 as follows:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

The nitrogen gas produced quickly inflates the bag between the driver and the windshield and dashboard. Calculate the volume of N_2 generated at 80 °C and 823 mmHg by the decomposition of 60g NaN_3?

T = 80 °C = 80 +273 = 353K

P= 823 mmHg

1 atm =760 mmHg

?atm = 823mmHg

823 x 1 = 760 x ? = 823/760 =1.083atm

m= 60g.

First convert g to mole

n=m/MM

= 60 / 65.02 =0.923 mol

From equation

2 mol NaN_3 → 3 mol N_2

0.923 mol NaN_3 → ? mol N_2

X

3 x 0.923 = 2 x ?

Mole of N_2= 3 x 0.923 / 2

Mole of N_2 = 1.38mol

Gas Stoichiometry

PV = nRT

$$V = \frac{n\,R\,T}{P}$$

$$V = \frac{1.38 \text{ x } 0.0821 \text{ x } 353}{1.083}$$

V = 36.9L

Dalton's Law of Partial Pressure

$$P_A = \frac{n_A RT}{V}$$

$$P_B = \frac{n_B RT}{V}$$

$$P_T = P_A + P_B$$

$$P_T = \frac{n_A RT}{V} + \frac{n_B RT}{V}$$

$$P_T = \frac{RT}{V}(n_A + n_B)$$

$$P_T = \frac{nRT}{V}$$

Where $n = n_A + n_B$

Dalton's Law of Partial Pressure

- Mole fraction: is a dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present.

$P_A = n_A RT/V$ Divided by P_T

$$\frac{P_A}{P_T} = \frac{n_A RT/V}{(n_A + n_B)RT/V}$$

$$\frac{P_A}{P_T} = \frac{n_A}{n_A + n_B}$$

$$= X_A$$

$$X_i = \frac{n_i}{n_T}$$

$$\frac{P_i}{P_T} = X_i$$

$$P_i = X_i P_T$$

If we have gas mixture consist of two gases (A and B)

Then the sum of all mole fraction for the same mixture is 1

$$X_A + X_B = \frac{n_A}{n_B + n_A} + \frac{n_B}{n_A + n_B} = 1$$

Dalton's Law of Partial Pressure

Example:

A mixture of gasses contains 4.46 moles of Ne, 0.74 mole of Ar, and 2.15 moles of Xe. Calculate the partial pressures of the gases if the total pressure is 2.00 atm at a certain temperature.?

First we have to determine the molar fraction of each gas

$$X_i = \frac{n_i}{n_T}$$

$$X_{Ne} = \frac{4.46}{4.46 + 0.74 + 2.15} = 0.607$$

$$X_{Ar} = \frac{0.74}{4.46 + 0.74 + 2.15} = 0.1$$

$$X_{Xe} = \frac{2.15}{4.46 + 0.74 + 2.15} = 0.293$$

$$P_{Ne} = X_{Ne} P_T = 0.607x2 = 1.214 atm$$

$$P_{Ar} = X_{Ar} P_T = 0.1x2 = 0.2 atm$$

$$P_{Xe} = X_{Xe} P_T = 0.293x2 = 0.586 atm$$

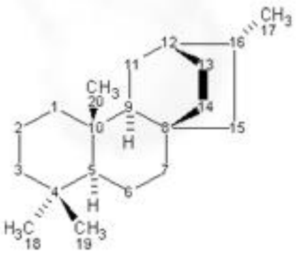
Examples

If the temperature and pressure are kept constant during the process, how many liters of $TiCl_4$ gas will be produced when 20.0 L of chlorine (Cl_2) react with titanium (Ti) according to the reaction: Ti(s) + 2 Cl_2 (g) → $TiCl_4$ (g.

a) 5.0 L
b) 10.0 L
c) 20.0 L
d) 40.0 L

From the equation
2 L of Cl2 produce 1 L of TiCl4
20 L of Cl2 ------------X L of TiCl4

X L of Cl2 = 10 L



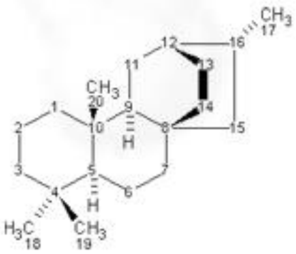
What is the pressure in atmospheres of a gas mixture that consists of 0.20 moles of nitrogen and 0.30 moles of oxygen in a 1250 mL container at 0 °C?

N moles of N2 = 0.20 mole
N moles of O2 = 0.30 mole
V = 1250 ml / 1000= 1.25 L
T= 0+ 273 = 273 K

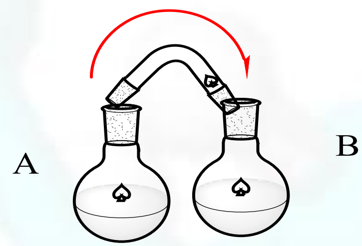
$$P_T = \frac{n\,R\,T}{V}$$

Where $n = n_A + n_B$

= (0.20 + 0.30) X 0.0821 X 273 / 1.25
= 8.95 atm



You have a sample of CO_2 gas in a flask (A) with a volume of 265 mL. At 22.5 °C, the pressure of the gas is 136.5 mmHg. To find the volume of another flask (B), you move the CO_2 to that flask and find that its pressure is now 94.3 mmHg at 24.5 °C. What is the volume of flask B?



a) 184.0 mL
b) 365.1 mL
c) 381.5 mL
d) 386.2 mL

V of A = 256 ml / 1000 = 0.256 L
T of A= 22.5 + 273 = 295.5 K
P of A= 136.5 mmHg / 760 = 0.179 atm

V of B = ?
T of B = 24.5 + 273 = 297.5 K
P of B = 94.3 mmHg / 760 = 0.124 atm

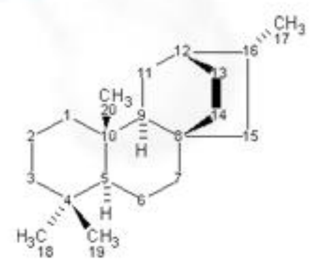
$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

A B

P1 V1 T2 = P2 V2 T1

V2 = P1V1T2 / P2 T1

= 386.2 mL



The pressure of 6.0 L of an ideal gas in a flexible container is decreased to one-third of its original value, and its absolute temperature is decreased by one-half. What is the final volume of the gas?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

a) 9.0 L
b) 6.0 L
c) 4.0 L
d) 1.0 L

V1 = 6 L

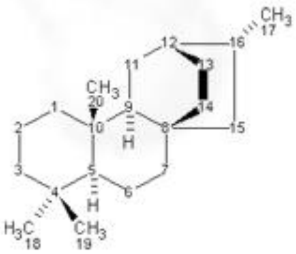
P2 = P1 /3

T2 = T1 /2

V2 = ?

$$V2 = \frac{P1\ \ V1\ \ T2}{P2\ \ T1}$$

$$V2 = \frac{P1 \times 6 \times \frac{T1}{2}}{-\frac{P1}{3} \times T1} = \frac{\frac{6}{2}}{\frac{1}{3}} = 9\ L$$



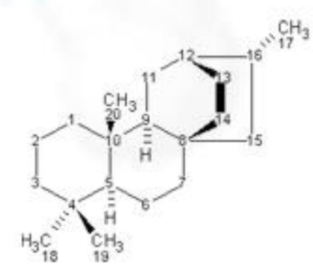
A fixed quantity of a gas is subjected to a decrease in pressure at constant temperature. The volume of the gas:

a) remains the same

b) Decreases

c) Increases

d) can’t be determined



A student adds 4.00 g of dry ice (solid CO_2) to an empty balloon. What will be the volume of the balloon at STP after all the dry ice sublimes (converts to gaseous CO_2)?

يجب ان نوجد عدد المولات

STP conditions means :
T = 0°C = 273.15 K
P = 1 atm.

$$n = 4 / 44 = 0.0909 \text{ mole}$$

PV = nRT

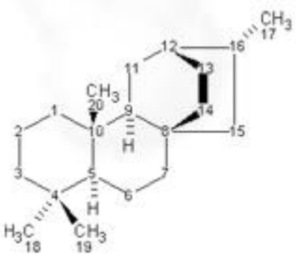
$$V = \frac{0.0909 \ X \ 0.0821 \ X \ 273.15}{1} = 2.038 \text{ L}$$

$$V = \frac{n\,R\,T}{P}$$

OR

At STP 1mole of an ideal gas occupies 22.414L.

V = n X 22.414
= 0.0909 X 22.414 = 2.037 L

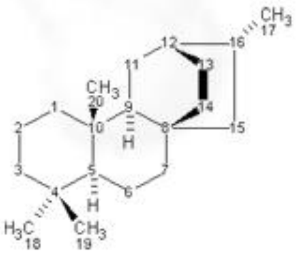


A compound is solid at room temperature, but it boils at 56 °C. Determine the density of the compound at 60 °C and 745 torr (molar mass of the compound = 352 g/mol).

$$d = \frac{P\,MM}{R\,T}$$

$T = 60 + 273 = 333$ K
$P = 745$ torr / $760 = 0.9802$ atm
$MM = 352$ g / mol

$$d = \frac{0.9802 \times 352}{0.0821 \times 333} = 12.64 \text{ g/L}$$



A mixture of two gases (A and B) are mixed in the same container. Calculate the mole fraction of gas B if the total pressure is 2 atm and the partial pressure of gas A is 1.5 atm ?

$P_T = 2$ atom

$P_A = 1.5$ atm

$X_B = ?$

$$X_A = \frac{P_A}{P_T}$$

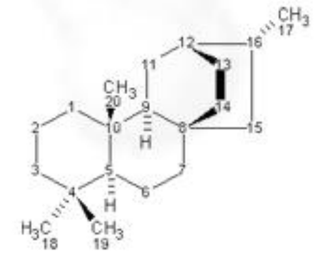
$$X_A = \frac{1.5}{2} = 0.75$$

$$X_A + X_B = 1$$

$$X_B = 1 - X_A$$

$$X_B = 1 - 0.75 = 0.25$$

$$X_A + X_B = \frac{n_A}{n_B + n_A} + \frac{n_B}{n_A + n_B} = 1$$



A certain amount of gas at 25 °C and at a pressure of 0.800 atm is contained in a glass vessel. Suppose that the vessel can withstand a pressure of 2.00 atm. How high can you raise the temperature of the gas without bursting the vessel?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

لدينا غاز تغيرت ضروفه من حاله لحاله لذلك نستخدم العلاقه

$T_1 = 25 + 273 = 298$ K

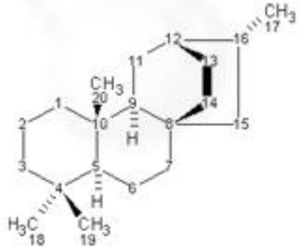
$$T_2 = \frac{P_2 V_2 T_1}{P_1 V_1} = \frac{P_2 T_1}{P_1}$$

$P_1 = 0.800$ atm

$P_2 = 2$ atm

$$T\,2 = \frac{2 \times 298}{0.800} = 745 \text{ K or } 472\ C^0$$

We assume that $V_2 = V_1$

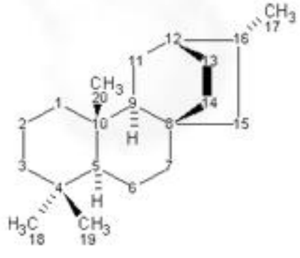


A sample of oxygen occupies 47.2 liters under a pressure of 1240 torr at 25$_o$C. What volume would it occupy at 25$_o$C if the pressure were decreased to 730 torr?

a) 27.8 L
b) 29.3 L
c) 32.3 L
d) **80.2 L**

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

اولا : حول درجة الحرارة الى كالفن وكذلك الضغط الى ضغط جوي ثم استخدم القانون اعلاه لحساب V_2 والاجابة الصحيحة باللون الاحمر



Under conditions of fixed temperature and amount of gas, Boyle's law requires that

I. $P_1V_1 = P_2V_2$

II. PV = constant

III. $P_1/P_2 = V_2/V_1$

نفرض ان ادينا القيم التالية من الضغط والحجم

a) I only

b) II only

c) III only

d) I, II, and III

P1 = 1 atm   P2 = 2 atm

V1 = 1 L   V2 = 0.50 L

I. $P_1V_1 = P_2 V_2$

1 X 1 = 2X 0.50

1 = 1

II. PV = constant

P1 x V1 = 1 x 1 = 1

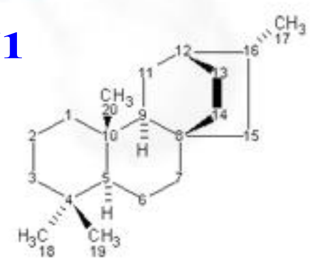
P2 x V2 = 2 x 0.50 = 1

III . P1 / P2 = V2 / V1

1/2 = 0.50 / 1

0.5 = 0.5

اذن نلاحظ ان العلاقات الثلاث صحيحة وبذلك تكون الاجابة الاخيرة باللون الاحمر هي الصحيحة



The volume of a sample of nitrogen is 6.00 liters at 35^{o}C and 740 torr. What volume will it occupy at STP?

a) 6.59 L
b) 5.46 L
c) 6.95 L
d) **5.18 L**

اولا نستخدم القانون

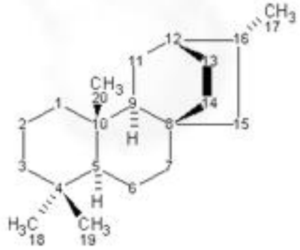
P V = n R T

لايجاد عدد المولات ثم بعد ذلك نضرب عدد المولات الناتجه في

22.414

الاجابة

5.18 L



What is the density of chlorine gas at STP, in grams per liter ?.

a) 6.2
b) **3.2**
c) 3.9
d) 4.5

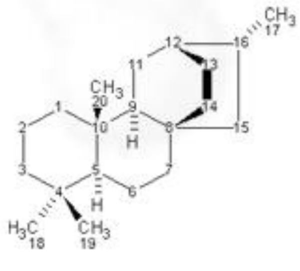
$$d = \frac{P\,MM}{R\,T}$$

نستخدم العلاقة

CL_2 = 2 x 35.5

STP means :
P = 1 atom
T= 273.15

$$d = \frac{1 \text{ X } 71}{0.0821 \text{ X } 273} = 3.2$$



What pressure (in atm) would be exerted by 76 g of fluorine gas in a 1.50 liter vessel at -37$_{o}$C?

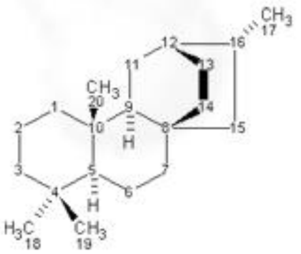
a) 26 atm
b) 4.1 atm
c) 19,600 atm
d) 84 atm

PV= nRT

P =

25.97 = 26 atm

V= 1.5 L
T= 273-37= 236 K



What is the density of ammonia gas at 2.00 atm pressure and a temperature of 25.0 ^{o}C?

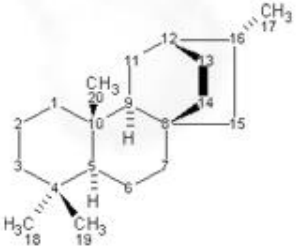
$$d = \frac{P\,MM}{R\,T}$$

a) 0.720 g/L
b) 0.980 g/L
c) **1.39 g/L**
d) 16.6 g/L

Convert 2.0 atm to mmHg

a) 150 mmHg
b) 0.27 mmHg
c) 150 mmHg
d) 1520 mmHg

1 atm = 760 mmHg

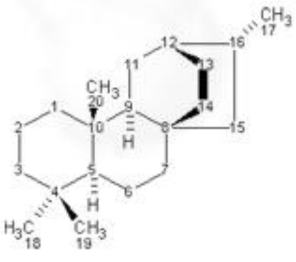


A container with volume 71.9 mL contains water vapor at a pressure of 10.4 atm and a temperature of 465_{o}C. How many grams of the gas are in the container?

a) 0.421 g
b) **0.222 g**
c) 0.183 g
d) 0.129 g

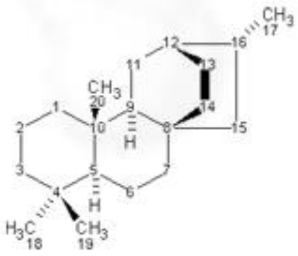
n = PV/RT = 0.0719× 10.4 = 0.0821 × (465 +273) = 0.012 mole

Mass = n × molar mass = 0.012 × 18 = 0.222 g



A 0.580 g sample of a compound containing only carbon and hydrogen contains 0.480 g of carbon and 0.100 g of hydrogen. At STP, 33.6 mL of the gas has a mass of 0.087 g. What is the molecular (true) formula for the compound?

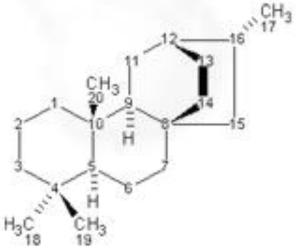
a) CH_3
b) C_2H_6
c) C_2H_5
d) C_4H_{10}



Gas occupy 6L at 37$_0$C what will be its volume when its temperature is doubled?

a) 12 L
b) 6L
c) 3.2 L
d) 2L

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$



A mixture of 90.0 grams of CH_4 and 10.0 grams of argon has a pressure of 250 torr under conditions of constant temperature and volume. The partial pressure of CH_4 in torr is:

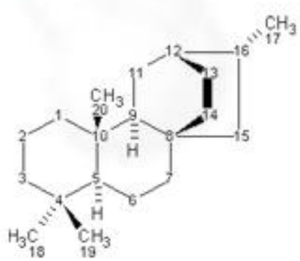
(a) 143
(b) 100
(c) 10.7
(d) 239

$n\ CH4 = 90 / 16 = 5.625$ mol
$n\ Ar = 10/39.9 = 0.250$ mol

$$X_{CH4} = 5.625 / 5.875 = 0.957$$

$$P_{CH4} = X_{CH4} \times P_T$$

$$0.957 \times 250 = 239 \text{ torr}$$



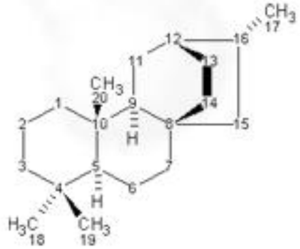
What pressure (in atm) would be exerted by a mixture of 1.4 g of nitrogen gas and 4.8 g of oxygen gas in a 200 mL container at 57_{o}C?

a) 4.7
b) 34
c) 47
d) 27

$P = n_{total} RT/V$

$n_{N2} = 1.4/ 2×14 = 0.05$ mole
$n_{O2} = 4.8/2×16 = 0.15$ mole

$P = (0.05 + 0.15)\ 0.0821×(57+273)\ /0.2 = 27$ atm



A sample of hydrogen gas collected by displacement of water occupied 30.0 mL at 24$_{o}$C and pressure 736 torr. What volume would the hydrogen occupy if it were dry and at STP? The vapor pressure of water at 24.0$_{o}$C is 22.4 torr

$P_{Total} = 736 / 760 = 0.968$ atm

$P_{H2O} = 22.4 / 760 = 0.029$ atm

$T = 24 + 273 = 297$ K

From dalton's law :

$P_{H2} = P_{total} - P_{H2O}$

$P_{H2} = 736 – 22.4 = 713.6$ torr

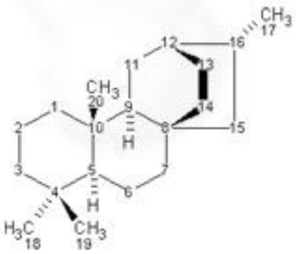
$n = PV / RT$

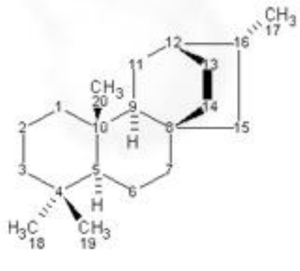
$n = (713.6/760) \times 0.03 / 0.0821 \times (24+273) = 0.00115$ mol

بعد ايجاد عدد مولات الهيدروجين نضربها في 22.4

V of H2 = 0.00115 x 22.4 = 0.026 L

V in ml = 0.026 X 1000 = 25.8 ml





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CH_3
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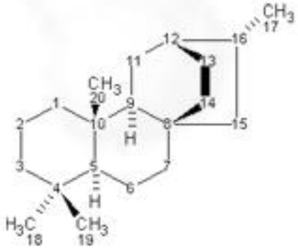

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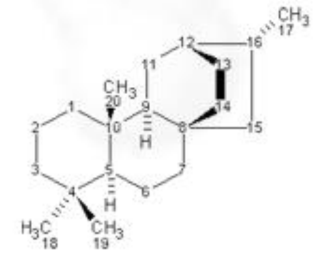

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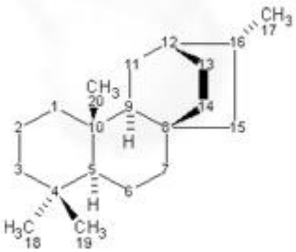

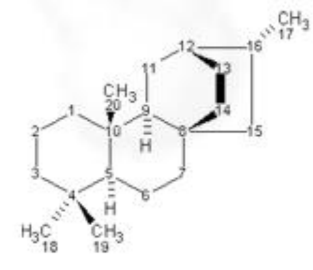

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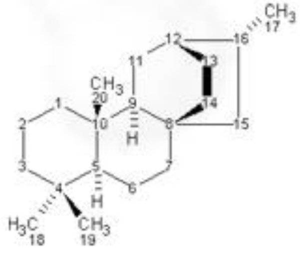
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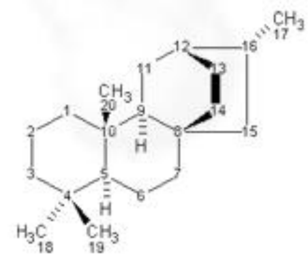

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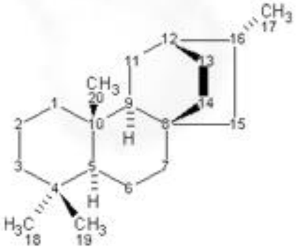

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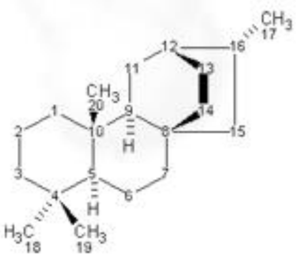
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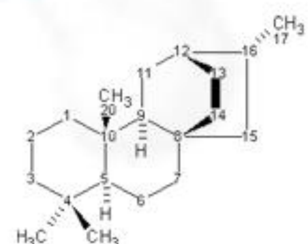

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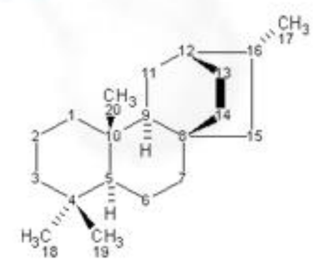
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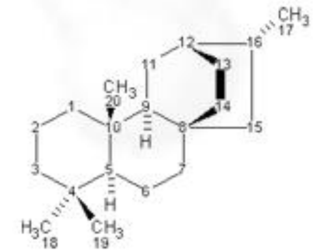

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Thank you for listening