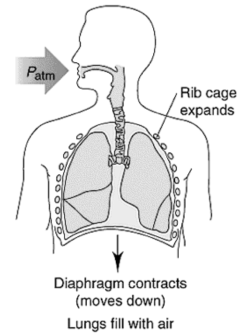
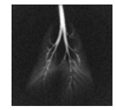
Gases

Chem 120 Prep

Dr. Ron Rusay





Do you have enough oxygen to climb Mt. Everest?

http://chemconnections.org/chemwiki/everest/everest.htm

QUESTION

Typical total volume for human lungs is approximately 5,800 mL. At a temperature of 37°C (average body temperature) and pressure of 0.98 atm, how many theoretical number of moles of air can we carry inside our lungs? (R = 0.08206 L atm/ K mol)

A) 1.9 mol
B) 0.22 mol
C) 230 mol
D) 2.20 mol
E) 0 mol: Moles can harm a person's lungs.

ANSWER

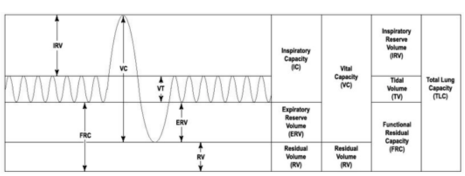
B)

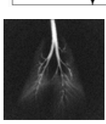
The units for temperature must be in K, pressure in atm, and volume in L. Then using the universal constant 0.08206 L atm/ K mol :

n_{air} = PV / RT

n_{air} = 0.98 atm x 5.800 L / (37 + 273) K x 0.08206 L * atm/ K mol

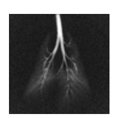
n_{air} = 0.22 mol





An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm and there is 20.9% oxygen in air. (R = 0.08206 L atm/ K mol)

1) *How many moles of O_2* are actually *in a typical breath?*.
2) *What is the mass of O_2 in a typical breath?*.
3) *How much of the O_2 is essential biochemically?*

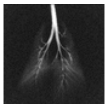


QUESTION

An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm and there is 20.9% oxygen in air. (R = 0.08206 L atm/ K mol)

How many moles of oxygen are actually *in a typical breath?*

A) 0.0020 mol
B) 0.020 mol
C) 0.030 mol
D) 0.025 mol
E) 0.0041 mol



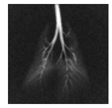
An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm and there is 20.9% oxygen in air.

How many moles of O_2 are actually *in a typical breath?*.

$n_{O_2(g)}$ = (20.9%) * PV / RT

$n_{O_2(g)}$ = (0.209 mol $O_{2(g)}$ / mol air) x 1.0 atm x (3.5 L-3.0 L) / 0.08206 L * atm x 310 K)

$n_{O_2(g)}$ = 0.0041 mol



ANSWER

An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm. (R = 0.08206 L atm/ K mol)

How many moles of air are actually *in a typical breath?*

A) 0.0020 mol
B) 0.020 mol air
C) 0.030 mol
D) 0.025 mol
E) 0.0041 mol oxygen

$n_{air\ (g)}$ = PV / RT

$n_{O_2(g)}$ = (20.9%) * PV / RT

QUESTION

An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm and there is 20.9% oxygen in air:

What is the mass of O_2 in a typical breath?

A) 0.0041 mol x 16 g/mol
B) 0.020 mol x 16 g/mol
C) 0.0041 mol x 32 g/mol
D) 0.020 mol x 32 g/mol

An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm and there is 20.9% oxygen in air.

What is the mass of O_2 in a typical breath?.

$n_{O_2(g)}$ = (20.9%) * PV / RT

$n_{O_2(g)}$ = (0.209 mol $O_{2(g)}$ / mol air) (1.0 atm x (3.5 L-3.0 L) x mol air * K / 0.0821 L * atm x 300 K)

$n_{O_2(g)}$ = 0.0041 mol

$g_{O_2(g)}$ = 0.0041 mol x 32.0 g/mol

$g_{O_2(g)}$ = 0.13 g

An average pair of human lungs actually contains only about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm

How much of the O_2 is essential biochemically?

Two estimates for a person with normal physical activity range from
0.67 - 0.84 kg of O_2 being used per day
(NASA provided the higher value). How many breaths do you take in one day? ~ 5 mol % of the O_2 is actually used per breath.

Hard exercise increases this oxygen demand (intake) about 10 fold.

QUESTION

The primary source of exhaled CO_2 is from the combustion of glucose, $C_6H_{12}O_6$ (molar mass = 180. g/mol.). The balanced equation is shown here:

$$C_6H_{12}O_6\ (aq) + 6\ O_2\ (g) \rightarrow 6\ CO_2\ (g) + 6\ H_2O\ (l)$$

If you oxidized 5.42 grams of $C_6H_{12}O_6$ while tying your boots to climb Mt. Everest, how many liters of O_2 @ STP conditions did you use? (R = 0.08206 L atm/ K mol)

A) 0.737 L
B) 0.672 L
C) 4.05 L
D) 22.4 L

ANSWER

C) 4.05 L

$C_6H_{12}O_6$ (*aq*) + **6** O_2 (*g*) → 6 CO_2 (*g*) + 6 H_2O (*l*)

The number of moles of glucose must first be determined (5.42 g /180. g/mol = 0.0301 moles), then this is multiplied by 6 to account for the stoichiometric ratio between glucose and oxygen.

From this, $V = nRT/P$ is used with the appropriate substitutions. (R = 0.08206 L atm/ K mol)

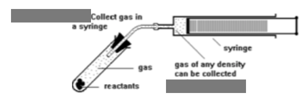
= **6 x** 0.0301 mol $\mathbf{O_2(g)}$ x 0.08206 L * atm mol^{-1} x 273 K/ 1 atm)

Molar Mass of a Gas

- ð *PV = n RT*
- ð *n = g of gas/ MM_{gas} [MM_{gas} = g/mol]*
- ð *PV = (g of gas/ MM_{gas})RT*
- ð *MM_{gas} = g of gas/V (RT/P)*

Density of gas = g of gas/V [experimental]

- ð *MM_{gas} = density of gas (RT/P)*

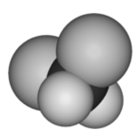


QUESTION

Freon-12, CF_2Cl_2, a "safe" compressible gas, was widely used from 1935-1994 as a refrigerant in refrigerators, freezers, and air conditioning systems. However, it had been shown to be a greenhouse gas and to catalytically destroy the ozone layer. It was phased out and banned.

200 ml of Freon-12 was collected by syringe. It weighed 0.927 grams, had a temperature of 30.0°C, and a pressure of 698.2 mm of Hg. What is the experimental molar mass of Freon-12? (R = 0.08206 L atm/ K mol)

A. 12.1 g/mol
B. 84 g/mol
C. 92.7 g/mol
D. 115 g/mol
E. 121 g/mol

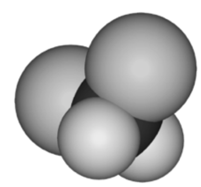


ANSWER

E)

MM_{gas} = density of gas (RT/P)

MM_{gas} = 0.927 g/ 0.200 L x 0.08206 L atm K^{-1} mol^{-1} x 303 K / 698.2 mm Hg / 760 mm Hg/ 1atm



QUESTION

The density of an unknown atmospheric gas pollutant was experimentally determined to be 1.964 g/ L @ 0 °C and 760 torr.

- *What is the molar mass of the gas?*
- *What might the gas be?*

A) CO B) SO_2 C) H_2O D) CO_2

ANSWER

CO (28g/mol) SO_2 (62g/mol) H_2O (18g/mol) CO_2 (44g/mol)

1.964 g/ L @ 0 °C and 760 torr.
R = 0.08206 L atm K^{-1} mol^{-1}
°C → K
torr → atm

MM_{gas} = density of gas (RT/P)
MM_{gas} = 1.964 g/ L x 0.08206 L atm K^{-1} mol^{-1} x 273K/ 760 torr x 760 torr/ 1atm

MM_{gas} = 44.0 g/mol

D) CO_2

QUESTION

0.0820 grams of the volatile, gaseous phase, of a compound, which smells like fresh raspberries, was trapped in a syringe. It had a volume of 12.2 mL at 1.00 atmosphere of pressure and 25.0°C. What is the molar mass of this pleasant smelling compound ?

A) 13.8 g/mol
B) 164 g/mol
C) 40.9 g/mol
D) 224 g/mol

ANSWER

B) 164 g/mol

Using $PV = nRT$:

0.0122 L for V, 298 K for T, 0.08206 for R and solving for n the number of moles represented by 0.0820 grams can be obtained.

Then the MOLAR MASS (grams in one mole) can be determined.

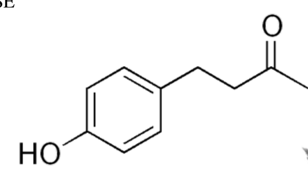
$$MM_{gas} = \text{density of gas } (RT/P)$$

$$MM_{gas} = 0.0820 \text{ g/ L} \times 0.00122 \text{ L} \times 0.0821 \text{ atm K}^{-1} \text{ mol}^{-1} \times 298\text{K}/ 1\text{atm}$$

QUESTION

For the compound that smells like fresh raspberries, the following structure, $C_{10}H_{12}O_2$, matches its calculated molar mass.

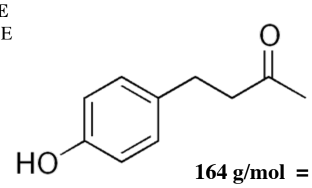
A) TRUE
B) FALSE



ANSWER

Based on your answers for the compound, which smells like fresh raspberries, in the previous two questions, the following structure matches its molecular formula.

A) TRUE
B) FALSE



164 g/mol = $C_{10}H_{12}O_2$