Gas Pressure and Volume

What is Pressure?

T minus 25 get ups

In studying the behaviour of gases we consider only "closed systems" (not open to atmosphere) like gas cylinders or balloons.

Gases exert pressure on all objects that they encounter. Pressure is the force exerted on an object per unit of surface area.

$\boxed{P = f/A}$

P = pressure
f = force
A = surface area

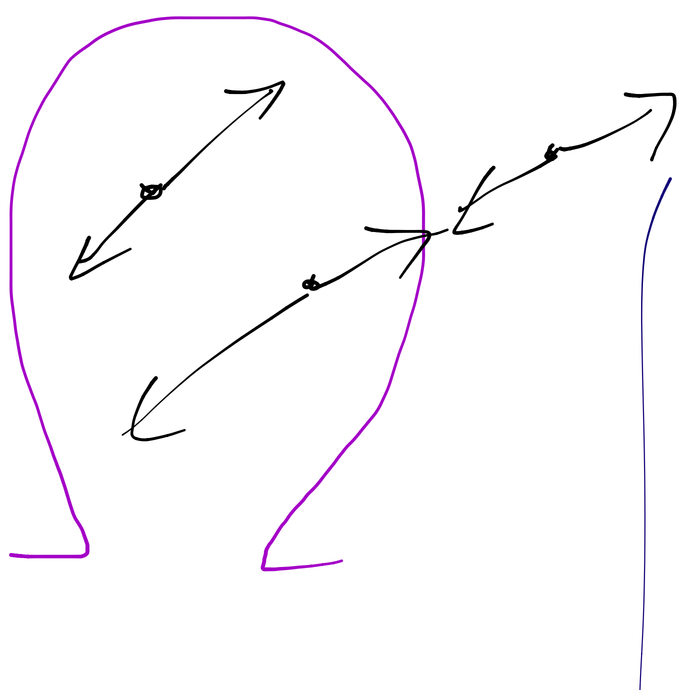
$P = F$

$P = \uparrow$ 87 times

The SI units for pressure is: pascals (Pa) or kilopasals (KPa)

In a container a gas exerts pressure because the molecules colliding with the walls of the container are collectively pushing out. The number of collisions (KE) determines the overall pressure.

Example: A balloon



There are a number of factors that can affect the number of collisions within a closed system. We will be looking at a few of these factors.

Pressure Measurements

Atmospheric pressure reflects the height of the column of air above us extending to the outermost layer of the atmosphere:

Troposphere, Stratosphere, Mesosphere, Thermosphere, Exosphere

Pressure can be measured in:

i) mm Hg 110/70

Represents the displacement of mercury.

Example: pressure of the atmosphere at sea level and 0 ºC (standard atmospheric pressure) is 760 mm Hg.

ii) torr

Represents a column of mercury 1 mm in height at 0 ºC

Example: standard atmospheric pressure at 0 ºC is 760 torr

iii) pascals (kilopascals)

Represents the force of 1 newton per metre squared. (1 N/m^2).

Example: standard atmospheric pressure at 0 ºC is equivalent to $\boxed{101.3\text{ kPa}}$.

iv) atmospheres (atm)

Represents the air pressure at sea level.

Example: standard atmospheric pressure at 0 ºC is $\boxed{1\text{ atm}}$ (760 torr)

All of the above units are equivalent, therefore:

CONVERSION FACTOR

Convert the following:

a) 3.22 atm to kPa

$$3.22\,atm \times \frac{101.3\,kPa}{1\,atm} = 326.2\,kPa$$

b) 945 mm Hg to kPa

c) 560 torr to atm

d) 112.3 kPa to torr

$$112.3\,kPa \times \frac{760\,torr}{101.3\,kPa} = 842\,torr$$

The problem with measuring a gas is that a gas has no fixed volume. Therefore you must consider the conditions in which the gas exists in a closed container. These conditions include:

i) Pressure

Determines how much gas is squeezed into a particular volume
Example: A gas cylinder

ii) Temperature

Affects the motion (KE) of the molecules.
Example: A cold tire vs a warm/hot tire

To calculate a quantity of gas the volume at known temperature and pressure must be specified.

Pressure Volume Relationships

Boyle investigated the relationship between pressure and volume of a gas using mercury in a glass tube.

Boyle discovered an inverse relationship between pressure and volume when temperature and number of moles of the gas were kept constant. This has become known as **Boyle's Law.**

Boyle's Law states that:

"The volume of a fixed mass of gas at constant" temperature varies inversely as the applied pressure.

Mathematically, this can be written as:

The equation can be re-written so that the proportionality sign in replaced by a constant:

$$\boxed{P_1 V_1 = P_2 V_2}$$

For changes in P and V the equation can be written as:

$$P \propto \frac{1}{V}$$

Example Questions:

What is the volume of 4.8 L of hydrogen gas if the pressure on it increases from 55 kPa to 127 kPa (T is constant)?

V_1 P_1 P_2 $V_2 = ?$

$$P_1 V_1 = P_2 V_2$$

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

$$= \frac{(55)(4.8)}{(127)}$$

$$V_2 = 2.1\,L$$

$\therefore$ V_2 is 2.1 L

A 16.0 L fire extinguisher contains 15.0 L of water and 1.00 L of compressed air. When in use, the extinguisher must expel the last bit of water at a pressure of 110 kPa. What should the original compressed air pressure in the extinguisher be (T is constant)?

$$P_1 V_1 = P_2 V_2$$

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

$$= \frac{(110\,kPa)(16)}{(1)}$$

$$= 1760\,kPa$$

$\therefore$ P_1 is 1760 kPa

Volume and Temperature Relationships

Charles investigated the relationship between temperature and volume of a gas.

Charles discovered a direct relationship between temperature and volume when pressure and number of moles of the gas were kept constant.

Charle's Law states, that: $V \propto T$

"the volume of a fixed quantity of gas at constant pressure is directly proportional to the temperature (in Kelvin)"

For changes in V and T the equation can be written as:

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

As previously mentioned, when using Charle's Law, temperature is reported in Kelvin. The starting point for the Kelvin scale is 0 K, which is called absolute zero. The accepted value for 0 K is –273.15°C.

In theory, at this temperature kinetic energy would be zero and the volume of the gas would also be zero.

To calculate temperature in Kelvin when given degree Celsius using the following equation:

$$\mathbf{T_K = °C + 273}$$

Example Questions:

If 50.0 cm^3 of gas in a syringe at 15.0 °C and the syringe's position is allowed to move outward against constant atmospheric pressure, calculate the new volume of the hot gas at 85.0 °C.

$V_2 = \frac{V_1 T_2}{T_1}$ 1

$= \frac{(50)(358)}{(288)}$

$T_1 = 15 + 273$
$= 288\,K$ 3

$T_2 = 85 + 273$
$= 358\,K$ 4

$V_2 = 62.2\,cm^3$ 2

$\therefore V_2$ is $62.2\,cm^3$

2/5

Determine the final volume of 20.0 L of a gas whose temperature changes from -73.0 °C to 327 °C if the pressure remains constant.

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

$= \frac{(20)(600)}{200}$

$T_1 = -73 + 273$
$= 200\,K$

$T_2 = 327 + 273$
$= 600\,K$

$\therefore V_2$ is $60\,L$

$= 60\,L$

Pressure and Temperature Relationships

Gay-Lussac discovered that the pressure (P) exerted by a gas is directly proportional to its temperature.

Gay-Lussac's Law states that: $P \propto T$

"the pressure of a fixed amount of gas, at constant volume is directly proportional to its Kelvin temperature".

For changes in P and T the equation can be written as: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Example Question:

A cylinder of chlorine gas is designed to withstand 50 atm of pressure. The pressure gauge reads 35.0 atm at 23°C. A fire causes the temperature in the storage room to increase to 85.5 °C. What will the pressure gauge read at this temperature?

$T_1 = 23 + 273 = 296\,K$

$T_2 = 85.5 + 273 = 358.5\,K$

$$P_2 = \frac{P_1 T_2}{T_1} = \frac{(35)(358.5)}{(296)} \doteq 42.4\ atm$$

$\therefore P_2$ is $42.4\ atm$

Combined Gas Law

When we examine all of the gas laws that we have learned so far, you will notice that some of the variables appear in multiple equations:

Gas Law	Variables	Constant	Equation
Boyles Law	P V	T	$P_1V_1 = P_2V_2$
Charles Law	V T	P	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay Lussac's Law	P T	V	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

We can combine all of the three laws to obtain the Combined Gas Law:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Before we look at some example problems using the Combined Gas Law, it is important we know a few of the standard conditions used when dealing with gases. A convenient way to express standard measurements conditions is to refer to STP or SATP conditions.

i) STP (standard temperature and pressure)
   T = 0°C or 273 K
   P = 760 mmHg, 760 torr, 1 atm or 101.3 kPa

ii) SATP (standard ambient temperature and pressure)
   T = 25 °C or 298 K
   P = 100 kPa

} I've seen this on the EXAM :)

Example Question:

A weather balloon contains 2.50 m^3 (V_1) of helium gas at 15.0 °C ($T_1 = 288K$) and 98.0 kPa ($P_1 = 98 kPa$). What is the volume at STP?

$T_2 = 273K$

$P_2 = 101.3 kPa$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{(98)(2.5)(273)}{(288)(101.3)}$$

$$V_2 \doteq 2.3 m^3$$

$V_2 = \frac{P_1 V_1 T_2}{T_1 P_2}$ 4 V_2 is $2.3\,m^3$

Molar Volume of Gases

Gay-Lussac measured the volumes of gases before and after a reaction. The research led him to devise the **law of combining volumes**:

"When gases react, the volumes of the reactants and products, measured at equal temperatures and pressures, are always in whole number ratios".

Example 1: hydrogen gas + oxygen gas → water vapour

$$2\,H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$
$$100\text{ ml} + 50\text{ ml} \rightarrow 100\text{ ml}$$

Example 2: ammonia gas → hydrogen gas + nitrogen gas

$$2NH_{3(g)} \rightarrow 3\,H_{2(g)} + N_{2(g)}$$
$$8\text{ ml} \rightarrow 12\text{ ml} + 4\text{ ml}$$

John Dalton examined the masses of compounds before and after a reaction, which led him to propose the law of multiple proportions:

"The masses of the elements that combine can be expressed in small whole number ratios".

By combining these ideas, **Amedeo Avogadro** related the volume of a gas to the amount that is present (calculated from the mass). Avogadro hypothesized that:

"Equal volumes of all ideal gases at the same temperature and pressure contain the same number of molecules".

Example 3: hydrogen gas + oxygen gas → water vapour

$$2\,H_{2(g)} + O_{2(g)} \rightarrow 2\,H_2O_{(g)}$$
$$2\text{ mol} + 1\text{ mol} \rightarrow 2\text{ mol}$$
$$2\text{ volumes} + 1\text{ volume} \rightarrow 2\text{ volumes}$$

$T = 273\,K$ $P = 1\,atm$

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

n = number of moles
V = Volume of gas

Example 4: At STP, 1 mol of oxygen gas has a volume of 22.4 L. Determine the mass and number of molecules in a 44. 8 L sample of the gas.

1 $n_2 = \frac{n_1 V_2}{V_1}$

$= \frac{(1)(44.8)}{22.4}$

3 $m = nM$

$= (2)(31.998)$

4 $= 63.996\,g$

$N = nN_A$

$\rightarrow N = 1.2 \times 10^{24}$ molecules

6   7

22.4

$\sqrt{2} \doteq 2\ mol$ $\quad$ $\sqrt{5}\ N = n\ N_A = (2)(6.02 \times 10^{23}))$

The **molar volume** of a gas is the space that is occupied by one mole of the gas. Molar volume is measured in units of **L/mol**, and is determined by dividing the volume by the number of moles that are present.

M $\boxed{v = \dfrac{V}{n}}$

v = molar volume (L/mol)
V = Volume of gas (L)
n = number of moles of gas (mol)

The molar volume of one gas is roughly the same as the molar volume of another gas at the same temperature and pressure. The molar volume of an ideal gas at STP is **22.4 L/mol**.

Example 5: An empty, sealed container has a volume of 0.652L (V_1) and a mass of 2.50 g. When filled with nitrogen gas, the container has a mass of 3.23 g. The pressure (P_1) of the nitrogen in the container is 97.5 kPa when the temperature is 21.0°C ($T_1 = 294\ K$). Calculate the molar volume of nitrogen gas at STP. $T_2 = 273\ K$ $\quad P_2 = 101.3\ kPa$

$$\sqrt{1}\ V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(97.5)(0.652)(273)}{(101.3)(294)}$$

$$\sqrt{2}\ = 0.58\ L$$

$$\sqrt{3}\ m_{N_2} = m_{container+N_2} - m_{cont}$$
$$= 3.23 - 2.5$$
$$\sqrt{6} = 0.73\ g$$

$$\sqrt{7}\ n_{N_2} = \frac{m}{M} = \frac{0.73}{28.014}$$
$$\sqrt{8} = 0.026\ mol$$

$$\sqrt{3}\ Mv = \frac{V}{n} = \frac{(0.58)}{0.026}$$
$$\sqrt{4} \doteq 22.3\ L/mol$$

$\sqrt{9}$ ∴ MV is 22.3 L/mol

The Ideal Gas Law and Stoichiometry

The Ideal Gas Law states that the pressure multiplied by the volume is equal to the number of moles multiplied by the universal gas constant and the temperature.

PV = nRT

$$R = 8.314 \frac{kPa \cdot L}{mol \cdot K}$$

Guidelines for using the Ideal Gas Law:

i) Convert temperature to Kelvin (K)
ii) Convert masses to moles (n)
iii) Convert volume to Litres (L)
iv) Convert pressure to kilopascals (kPa)

Sample Problems

1. 2.00 L of nitrogen dioxide gas in a container holds 4.2 mol at 206 kPa. What is the temperature in the cylinder?

$$T = \frac{PV}{nR} = \frac{(206)(2)}{(4.2)(8.314)}$$

$T = 11.8\,K$

$\therefore$ T is 11.8K

2. Calculate the volume that 6.30 mol of carbon dioxide gas at 23 ºC and 550 kPa pressure occupy.

$T = 23 + 273 = 296\,K$

$$V = \frac{nRT}{P} = \frac{(6.3)(8.314)(296)}{(550)} = 28.2\,L$$

$\therefore$ V is 28.2 L

3. 1.5 L of propane gas is burned in a barbeques. The following equation shows the reaction. Assume all gases are at STP:

$$\_\_\_ C_3H_{8\,(g)} + \underline{5}\, O_{2\,(g)} \rightarrow \underline{3}\, CO_{2\,(g)} + \underline{4}\, H_2O_{(g)}$$

$V = 1.5\,L$ $V = ?$

What volume of carbon dioxide is produced?

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

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

$$V_2 = \frac{(3)(1.5)}{(1)} = 4.5\,L$$

$\therefore$ V is 4.5 L

Dalton's Law of Partial Pressure

Consider a container that contains three types of gases:
N_2, O_2, CO_2

The molecules of each gas collide with the container walls and with the sensor on the pressure guage.

Each gas exerts a pressure called its **partial pressure**.

The total pressure is therefore the sum of all the partial pressures of the gases present.

As an equation, this can be written as:

$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2}$$

Example:

A closed container contains a mixture of O_2 and CO_2. If the pressure is 2.68 atm and the temperature is 273 K, calculate the partial pressure of O_2 if the mixture is 30% CO_2. Convert your answer to kPa.

$$\%CO_2 = \frac{P_{CO_2}}{P_{total}} \times 100$$

1 $$P_{CO_2} = \frac{(\%CO_2)(P_{total})}{100}$$

$$= \frac{(30)(2.68)}{100}$$

2 $$= 0.804 \text{ atm}$$

3 $$P_{O_2} = P_{total} - P_{CO_2}$$

$$= 2.68 - 0.804$$

4 $$= 1.876 \text{ atm}$$

5 $$P_{O_2} = 1.876 \times \frac{101.3}{1}$$

6 $$\therefore = 190 \text{ kPa}$$

7 $\therefore$ P_{O_2} is 190 kPa

Gas Applications

Finding the Density of a Gas

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

Oxygen gas makes up about 20% of our atmosphere. Find the density of pure oxygen gas in g/L at 12.50°C and 126.63 kPa.

$PV = nRT$

$T = 273 + 12.5$

$T = 285.5\,K$

2 Assume 1 L of $O_{2(g)}$

3 $n = \frac{PV}{RT}$

$= \frac{(126.63)(1)}{(8.314)(285.5)}$

4 $\doteq 0.053$ mol

5 $m = nM$

$= (0.053)(31.998)$

6 $\doteq 1.71$ g

7 $D = \frac{m}{V}$

$= \frac{1.71}{1}$

8 $\doteq 1.71$ g/L

9 $\therefore O_2$ is 1.71 g/L

Using Molar Mass to Identify an Unknown Gas

A 1.56 L gas sample has a mass of 3.22 g at 100 kPa and 281 K. What is the molar mass of the gas?

$M = \frac{m}{n}$

$PV = nRT$

$n = \frac{PV}{RT}$

$= \frac{(100)(1.56)}{(8.314)(281)}$

$\doteq 0.07$ mol

$M = \frac{m}{n}$

$= \frac{3.22\,g}{0.07\,mol}$

$= 48.2$ g/mol