Behaviour of gases

CHANGING THE VARIABLES FOR A FIXED MASS OF GAS

$P \propto \frac{1}{V}$ (or PV = constant)

At constant temperature: as the volume decreases the concentration of the particles increases, resulting in more collisions with the container walls. This increase in pressure is inversely proportional to the volume, i.e. doubling the pressure halves the volume.

$V \propto T$ (or $\frac{V}{T}$ = constant)

At constant pressure: at higher temperatures the particles have a greater average velocity so individual particles will collide with the container walls with greater force. To keep the pressure constant there must be fewer collisions per unit area so the volume of the gas must increase. The increase in volume is directly proportional to the absolute temperature, i.e. doubling the absolute temperature doubles the volume.

$P \propto T$ (or $\frac{P}{T}$ = constant)

At constant volume: increasing the temperature increases the average kinetic energy so the force with which the particles collide with the container walls increases. Hence pressure increases and is directly proportional to the absolute temperature, i.e. doubling the absolute temperature doubles the pressure.

Real gases

An ideal gas exactly obeys the gas laws. As real gases do have some attractive forces between the particles and the particles themselves do occupy some space so they do not exactly obey the laws. If they did they could never condense into liquids. A gas behaves most like an ideal gas at high temperatures and low pressures.

IDEAL GAS EQUATION

The different variables for a gas are all related by the ideal gas equation.

$$PV = nRT$$

P = pressure in Pa (N m^{-2})
(1 atm = 1.013×10^5 Pa)
T = absolute temperature in K
V = volume in m^3
(1 cm^3 = 1×10^{-6} m^3)
n = number of moles
R = gas constant = 8.314 J K^{-1} mol^{-1}

Units

The gas constant can be expressed in different units but it is easier to use SI units.

$$R = \frac{PV}{nT} = \frac{\text{N m}^{-2} \times \text{m}^3}{\text{mol} \times \text{K}} = \text{N m mol}^{-1}\ \text{K}^{-1}$$

$$= \text{J K}^{-1}\ \text{mol}^{-1}$$

Molar volume of a gas

The ideal gas equation depends on the amount of gas (number of moles of gas) but not on the nature of the gas. One mole of any gas will occupy the same volume at the same temperature and pressure. At 273 K and 1.013×10^5 Pa (1 atm) pressure this volume is 2.24×10^{-2} m^3 (22.4 dm^3 or 22 400 cm^3).

When the mass of a particular gas is fixed (nR is constant) a useful expression to convert the pressure, temperature, and volume under one set of conditions (1) to another set of conditions (2) is:

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

In this expression there is no need to convert to SI units as long as the same units for pressure and volume are used on both sides of the equation. However, do not forget that T refers to the absolute temperature and must be in kelvin.

Worked example 1

What volume will be occupied by 0.216 g of carbon dioxide at 21 °C and at a pressure of 1.32 atm?

Step 1. Calculate the number of moles of gas

$$\text{Amount of carbon dioxide} = \frac{0.216}{44.0} = 4.91 \times 10^{-3}\ \text{mol}$$

Step 2. Express all temperatures as absolute temperatures

21 °C = 294 K

Step 3. Convert all other units to SI units

1.32 atm = $1.32 \times 1.013 \times 10^5 = 1.34 \times 10^5$ Pa

Step 4. Apply ideal gas equation $PV = nRT$

$$1.34 \times 10^5 \times V = 4.91 \times 10^{-3} \times 8.314 \times 294$$

$$V = \frac{4.91 \times 10^{-3} \times 8.314 \times 294}{1.34 \times 10^5} = 8.96 \times 10^{-5}\ \text{m}^3\ (89.6\ \text{cm}^3)$$

Worked example 2

A gas occupies 127 cm^3 at a pressure of 0.830 atm and at 28 °C.

(a) What volume will the same amount of gas occupy at 1.00 atm pressure and at 0 °C?

(b) How many moles of gas are present?

(a) Step 1. Express all temperatures as absolute temperatures

28 °C = 301 K   0 °C = 273 K

Step 2. Apply $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

(Note that there is no need to change the units to SI units.)

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} = 127 \times \frac{0.830}{1.00} \times \frac{273}{301} = 95.6\ \text{cm}^3$$

(b) The number of moles can be calculated by using either the molar volume of a gas **(i)** or the ideal gas equation **(ii)**.

(i) 1 mole of any gas occupies 22 400 cm^3 at 273 K, 1 atm

Amount of gas occupying 95.6 cm^3 at 273 K, 1 atm $= \frac{95.6}{22\,400} = 4.27 \times 10^{-3}$ mol.

(ii) Convert all units to SI units:

127 cm^3 = 1.27×10^{-4} m^3
0.830 atm = $0.830 \times 1.013 \times 10^5$
= 8.41×10^4 Pa

$$n = \frac{PV}{RT} = \frac{8.41 \times 10^4 \times 1.27 \times 10^{-4}}{8.314 \times 301}$$

$$= 4.27 \times 10^{-3}\ \text{mol}$$