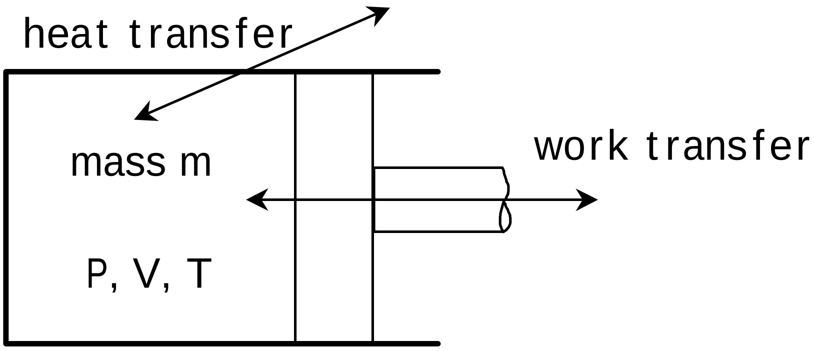
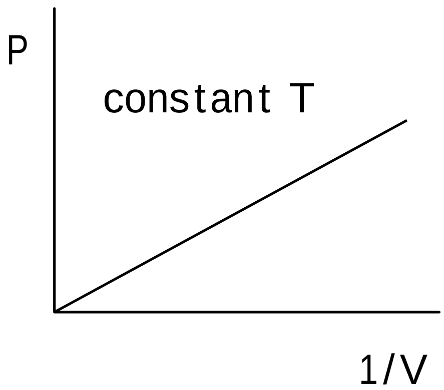
1.5 The Gas Laws

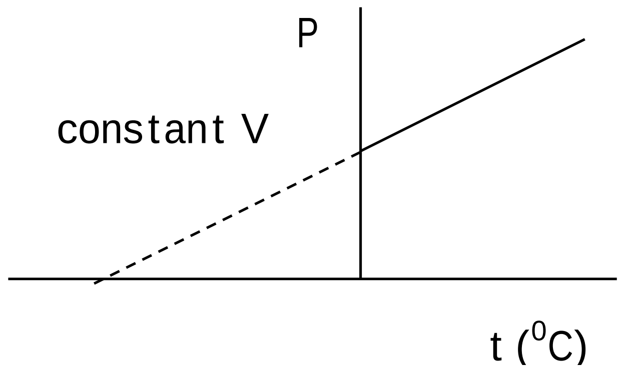
Early experiments on gases showed that the pressure, volume, and temperature of a gas were related:



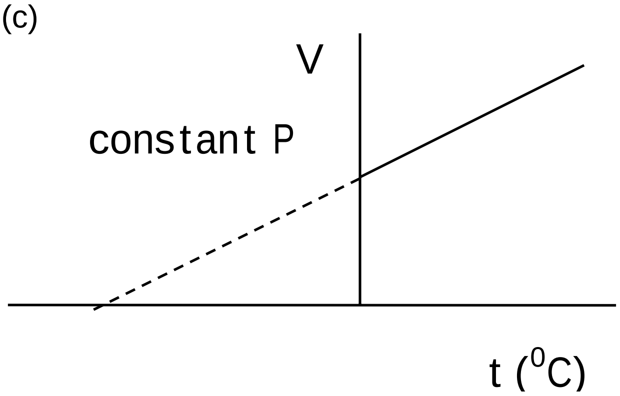
(a)



(b)



(c)



(a) Boyle showed that if the temperature was kept constant, then the pressure was inversely proportional to the volume.

How would you write this mathematically?

(b) and (c) If the volume was kept constant, then the pressure was directly proportional to the temperature, while if the pressure was kept constant, then the volume was directly proportional to the temperature.

Experimenters found that if the lines on the p-T and V-T graphs were extended backwards, then, in theory, a gas would have zero pressure, or zero volume at –273.15°C.

What is this temperature now known as?

Putting all these results together, the "ideal gas equation" or perfect gas model was formulated:

$$pV = mRT$$

where
- p = pressure in Pa
- V = volume in m^3
- m = mass in kg
- R = the gas constant (for the particular gas) in J kg^{-1} K^{-1}
- T = the absolute temperature in Kelvin

or dividing both sides by m,

$pv = RT$ where v is the specific volume (v = V/m) in m^3 kg^{-1}

Alternatively, $\frac{p}{\rho} = RT$ where ρ is the density (ρ = 1/v) in kg m^{-3}

This is called the **equation of state**, for it relates the properties of an ideal gas to one another, in a particular state.

For the properties of a fixed mass of gas to change, a **process** must occur, which will mean a **transfer of energy** to or from the gas, by means of heat transfer or work transfer. The process brings about a change of state, from state 1 to state 2. This means that for a fixed mass of a particular gas,

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

So at constant temperature, i.e. $T_1 = T_2$, then $p_1V_1 = p_2V_2$.

More generally, Boyle's Law is expressed as pV = constant at constant temperature.

The temperature, -273.15°C is absolute zero temperature, 0 K.

Remember that when applying the equation of state, the temperature is always the absolute temperature, expressed in Kelvin.

To convert temperatures from Celsius to Kelvin,

$$T(K) = t(°C) + 273.15$$

In general, the **gas constant** $R = \frac{R_0}{MM}$

where R_0 is the **universal gas constant**, 8314. J kmol^{-1} K^{-1}
MM is the **relative molecular mass** of the gas (equals its molecular weight)

Thus each gas has its own unique value of R. For air, R = 287 J kg^{-1} K^{-1}.

Example

A gas that can be modelled as perfect has a molecular mass of 44. A mass of 0.01 kg is enclosed in a container of volume 0.005 m^3 at a temperature of 17°C. (a) What is its density under these conditions? (b) What must its pressure be?

(a) m = 0.01 kg V = 0.00 m^3

Density, $\rho = \frac{m}{V} = \frac{0.01}{0.005} = 2 \text{ kg m}^{-3}$

(b) MM = 44 m = 0.01 kg V = 0.005 m^3 t = 17°C

For this gas, $R = \frac{R_0}{MM} = \frac{8314.3}{44} = 189 \text{ J kg}^{-1}\text{K}^{-1}$

The absolute temperature T = t +273.15 = 17 + 273.15 = 290.15 K

$pV = mRT \;\; \therefore p = \frac{mRT}{V} = \frac{0.01 * 189 * 290.15}{0.005} = 109676.7 \text{ Pa}$

The pressure of the gas is 109.7 kPa

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