InkPhysics

ANSWER BOOK

Molecular kinetic theory

Thermal physics · Year 13

AQA 3.6.2.3 · CIE 15.3.1, 15.3.2, 15.3.3, 15.3.4
OCR A 5.1.2(c), 5.1.4(b), 5.1.4(c), 5.1.4(e), 5.1.4(f), 5.1.4(h), 5.1.4(i)

19 questions · 53 marks

NAVY EDITION

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SECTION A · Warm-up

A1 Any two of the following, one mark each (2): molecules are in constant random motion; collisions with the walls and each other are perfectly elastic; the volume of the molecules is negligible compared with the container; the time of collisions is negligible; and there are no forces between molecules except during collisions. [2]

A2 Mean KE = (3/2)kT = 1.5 × 1.38 × 10^{-23} × 300 (1)
Mean KE = 6.21 × 10^{-21} J (1) [2]

A3 $pV = \frac{1}{3}Nm(c_{rms})^2$ (1), where N is the number of molecules, m the mass of one molecule and c_{rms} the root-mean-square speed (1). [2]

A4 The smoke particles move continually in random, jerky paths (Brownian motion) (1). This is evidence that air consists of molecules in rapid random motion, bombarding the smoke particles unevenly (1). [2]

A5 The square root of the mean of the squares of the molecular speeds (1). [1]

A6 Mean square speed = $(300^2 + 400^2 + 500^2 + 600^2)/4 = 2.15 \times 10^5$ m^2 s^{-2} (1)
$c_{rms} = \sqrt{(2.15 \times 10^5)} = 464$ m s^{-1} (1) [2]

SECTION B · Standard

B1 Each molecule carries (3/2)kT on average, so a mole carries (3/2)RT (1)
= 1.5 × 8.31 × 400 (1)
= 4986 J ≈ 5.0 × 10^3 J (1) [3]

B2 $c_{rms} = \sqrt{(3kT/m)}$ (1)
$= \sqrt{((3 \times 1.38 \times 10^{-23} \times 300)/(4.7 \times 10^{-26}))}$ (1)
$c_{rms} = 514$ m s^{-1} (1) [3]

B3 $p = \frac{1}{3}\rho(c_{rms})^2$ (1)
$= (1/3) \times 1.2 \times 500^2$ (1)
p = 1.00 × 10^5 Pa (1) [3]

B4 The light molecule (1). At a given temperature all molecules have the same average kinetic energy (3/2)kT, so $\frac{1}{2}m(c_{rms})^2$ is the same (1); a smaller mass m therefore requires a larger c_{rms}, since $c_{rms} \propto 1/\sqrt{m}$ (1). [3]

B5 [3]

$\Delta p = 2mv = 2 \times 4.65 \times 10^{-26} \times 480$ (1)
$\Delta p = 4.46 \times 10^{-23}$ kg m s^{-1} (1)
$F = \Delta p/t = 4.46 \times 10^{-23}/(1.2 \times 10^{-3}) = 3.7 \times 10^{-20}$ N (1)

B6 [3]

The mean kinetic energies are equal (ratio 1), because mean KE = (3/2)kT depends only on temperature (1)
Equal $\frac{1}{2}m(c_{rms})^2$ gives $c_{rms} \propto 1/\sqrt{m}$, so the ratio = $\sqrt{(m_{Ar}/m_{He})} = \sqrt{10}$ (1)
Ratio = 3.2 (1)

B7 [2]

Collisions of molecules with the walls (and with each other) are perfectly elastic, so no kinetic energy is lost in collisions (1). The molecules therefore keep the same mean square speed, the rate of change of momentum at the walls is unchanged, and the pressure stays constant (1).

SECTION C · Stretch

C1 [4]

(a) Mean KE = (3/2)kT = $1.5 \times 1.38 \times 10^{-23} \times 290$ (1)
Mean KE = 6.0×10^{-21} J (1)
(b) $c_{rms} = \sqrt{(3kT/m)} = \sqrt{((3 \times 1.38 \times 10^{-23} \times 290)/(5.3 \times 10^{-26}))}$ (1)
c_{rms} = 476 m s^{-1} (1)

C2 [3]

$c_{rms} \propto \sqrt{T}$ (1)
Doubling c_{rms} therefore needs four times the absolute temperature (1)
T = 4 × 300 = 1200 K (1)

C3 [3]

Molecules move randomly and repeatedly collide with the walls (1). Each collision changes a molecule's momentum, so by Newton's laws the wall exerts a force on it and it exerts an equal force back (1). The many collisions per second over the wall area produce a steady average pressure (1).

C4 [5]

One molecule moving at velocity component v_x towards a wall rebounds elastically, a change in momentum of $2mv_x$ (1). It returns after travelling 2L, so it strikes the wall every $2L/v_x$ seconds, giving an average force $2mv_x/(2L/v_x) = mv_x^2/L$ (1). Summing over all N molecules and replacing v_x^2 by its mean gives a total force $Nm(v_x^2)_{mean}/L$ (1). The motion is random, so the mean square speed is shared equally among three directions: $(v_x^2)_{mean} = \frac{1}{3}(c_{rms})^2$ (1). Pressure = force/L^2 and $V = L^3$, so $pV = \frac{1}{3}Nm(c_{rms})^2$ (1).

C5 [4]

From $\frac{1}{2}m(c_{rms})^2 = \frac{3}{2}kT$: $T = m(c_{rms})^2/(3k)$ (1)

$T = 3.3 \times 10^{-27} \times (1.1 \times 10^4)^2/(3 \times 1.38 \times 10^{-23}) = 9.6 \times 10^3$ K (1)

The molecular speeds are spread about c_{rms}, so at a few hundred kelvin a small fraction of hydrogen molecules still move faster than the escape speed (1)

These fast molecules escape, and over a long time the hydrogen is lost (1)

C6 [3]

An ideal gas has no intermolecular forces (except in collisions), so there is no potential-energy contribution to the internal energy (1)

Total kinetic energy = $N \times \frac{3}{2}kT = \frac{3}{2}pV$, using $pV = NkT$ (1)

$U = 1.5 \times 2.5 \times 10^5 \times 0.020 = 7500$ J (1)