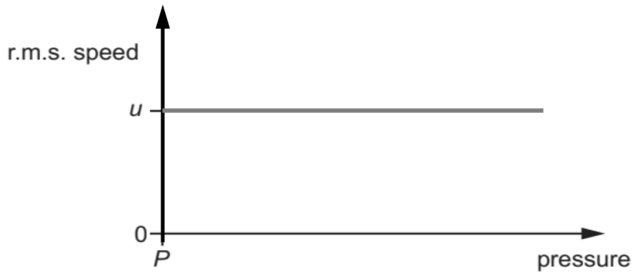
**1**

**(a)** **(i)** total volume of molecules negligible compared to that of containing vessel
time of collision small compared with the time between collisions

**(ii)** random kinetic energy increases with temperature
no potential energy
so increase in temperature increases internal energy

**(iii)**




**(b)** gas expands so work done by gas against atmosphere
negligible time for thermal energy Q to enter or leave the gas
Since $\Delta U = Q_{to} + W_{on}$ gas and hence $\Delta U$ is negative) $U$ decreases

**2**

**(b)** **(i)** $\frac{PV}{T} = nR$

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

$\frac{(101000)(15000)}{288} = \frac{890V_2}{228}$

$V_2 = 1.35 \times 10^6\ m^3$

**(ii)** $KE_{avg} = \frac{3}{2}kT$

$= \frac{3}{2}(1.38 \times 10^{-23})(228)$

$= 4.72 \times 10^{-21}\ J$

**(iii)** $\Delta U = U_f - U_i$

$= \frac{3}{2}\left(P_fV_f - P_iV_i\right)$

$= \frac{3}{2}\left(890 \times 1.3476 \times 10^6 - 101000(15000)\right)$

$= -4.73 \times 10^8\ J$

**3**

**(a)** gas that obeys $pV \propto T$ for all values of p, V and T
where p is pressure, V is volume and T is thermodynamic temperature

**(b)** $n = \frac{pV}{RT}$

$= \frac{(1.6 \times 10^6)(0.20)}{8.31(22 + 273.15)}$

$= 130\ mol$

**(c)** $\frac{1}{2}Mc_{rms}^2 = \frac{3}{2}RT$

$$c_{rms} = \sqrt{\frac{3RT}{M}}$$
$$= \sqrt{\frac{3(8.31)(22 + 273.15)}{4.2 \times 10^{-2}}}$$
$$= 420\ m\ s^{-1}$$

**(d)** When gas reaches equilibrium with surroundings, its pressure is $3.6 \times 10^4\ Pa$.

$$n' = \frac{p'V}{RT'} = \frac{(3.6 \times 10^4)(0.20)}{8.31(-50 + 273.15)} = 3.8827\ mol$$
$$m = 3.8827(4.2 \times 10^{-2}) = 0.16\ kg$$

**4**

**(a)** **(i)** Molecules collide with one another in a haphazard manner hence they possess different kinetic energies at any time.

**(iii)** No intermolecular forces in ideal gas, hence no potential energy

**(iii)**
$$p = \frac{1}{3}\rho < c^2 >$$
$$pV = \frac{1}{3}M < c^2 > since\ p = M/V$$
For ideal gas, $pV = nRT$
Hence $\frac{1}{3}M < c^2 >= nRT$
Total KE $= \frac{1}{2}M < c^2 >= \frac{3}{2}nRT$

**(iv)** Low pressure and high temperature

**5**

**(a)** An ideal gas obeys the equation pV = nRT where p is the gas pressure, V is the volume of the gas, n is the number of moles of gas, R is the universal gas constant, and T is the thermodynamic temperature of the gas. There are no intermolecular forces between the gas molecules, therefore an ideal gas has no potential energy.

**(b)** **(i)**
$$PV = NkT$$
$$N = \frac{PV}{kT}$$
$$= \frac{4.8 \times 10^5 \times 2.3 \times 10^4}{100^3(65 + 273.15) \times 1.38 \times 10^{-23}}$$
$$= 2.37 \times 10^{24}$$

**(ii)**
$$E_K = \frac{3}{2}NkT$$
$$= \frac{3}{2}(1.38 \times 10^{-23})(65 + 273.15)(2.37 \times 10^{24})$$
$$= 1.66 \times 10^4\ J$$

**(iii)**
$$\frac{1}{2}m < c^2 >= \frac{3}{2}kT$$
$$< c^2 > \propto T$$

$$\frac{c_{rms,65}}{c_{rms,75}} = \sqrt{\frac{65 + 273.15}{75 + 273.15}}$$
$$= 0.986$$

**6**

**(a)** **(i)** Gas atoms are in constant random motions, and they continually collide with the walls of the cylinder. When a gas atom collides with the wall of the cylinder, it rebounds and its velocity changes direction, hence there is a change in momentum of the gas atom.
By Newton's second law of motion, there is a resultant force exerted by the wall on the gas atom which is proportional to the rate of change in momentum.

By Newton's third law, the gas atom exerts a force of equal magnitude and opposite direction on the wall. The pressure exerted by the gas is the total force per unit area of the container walls exerted by all the gas atoms on the walls of its container.

**(ii)** The gas atoms are in random motion and this means that every atom has equal probability of moving in every direction. Also, the mean speed in all directions are equal. Consequently, the vector sum of all their velocities is equal to zero. Hence, the mean velocity is zero.

**7**

**(a)** **(i)** Random motion of a large number of molecules in random velocity means no preferred direction of movement or the molecules move in different/all directions with a root-mean square speed. Momentum is a vector given by the product of mass and velocity. Since the velocities cancel in opposite directions, the average velocity is zero and hence the average momentum of the gas molecules of same mass in a container is zero.

**(ii)** When the gas molecules hit the wall of the container, they rebound off the wall with the same speed and experiences a rate of momentum change normal to the wall, implying there is force acting on the molecule by the wall. Other molecules also have a momentum change normal to the wall when they strike and rebound.

By Newton's third law, the molecules exert an equal and opposite force on the walls. Since there are a large number of molecules, the number of collisions at any instant in time is very large and practically constant, resulting in a constant gas pressure. The pressure on the walls of the container is the average force per unit area exerted by the molecules on the walls of the container.

**8**

**(a)** **(i)** Kinetic energy of one gas particle (atom / molecule) $= \frac{1}{2}mc_{rms}^2 = \frac{3}{2}kT$
where $m$ is the mass of one gas particle/atom/molecule.

For one mole of gas containing $N_A$ particles, the total kinetic energy is given by:
$$\frac{1}{2}N_A mc_{rms}^2 = \frac{3}{2}N_A kT$$

From the equation of state of an ideal gas for 1 mole of ideal gas:
$$pV = (1)RT = N_A kT$$

$$\frac{1}{2}N_A mc_{rms}^2 = \frac{3}{2}RT$$

$$c_{rms} = \sqrt{\frac{3RT}{N_A m}} = \sqrt{\frac{3RT}{M}}$$

Where $N_A m$ is the mass of $6.02 \times 10^{23}$ particles = molar mass $M$.

**(ii)** $c_{rms} \propto \frac{1}{\sqrt{M}}$

$$\frac{c_{rms}\ of\ oxygen\ molecules}{c_{rms}\ of\ nitrogen\ molecules} = \sqrt{\frac{28}{32}} = 0.935$$

**(b)** **(i)** $P = \frac{1}{3}\left(\frac{NM}{V}\right) < c^2 >= \frac{1}{3} P < c^2 >= \frac{1}{3} \times 1.50 \times 10^5 \times (4.85 \times 10^5)^2$
$= 1.18 \times 10^{16}\ Pa$

**(ii)** Volume of particles is not negligible at that density. This results in a higher rate of collision compared to the expected rate, causing the actual pressure to be higher than the expected value in (b)(i).

**9**

**(a)** momentum after elastic collision $= -mu$
time between collisions $= \frac{2L}{u}$
number of collisions per unit time $= \frac{u}{2L}$
rate of change of momentum $= -\frac{mu^2}{L}$
average force $= \frac{mu^2}{L}$

**(b)** **(i)** Gas molecules collide with one another elastically.

**(ii)** Since collisions are elastic, the average speed of gas molecules incident on the wall remains the same. So frequency of collision between each wall and the molecule also remains unchanged. Therefore the pressure exerted on the wall is not affected.

**(c)**
$pV = \frac{1}{3} NM < c^2 >$
$pV = NkT$
$\frac{1}{3} NM < c^2 >= NkT$
$\frac{1}{2} m < c^2 >= \frac{3}{2} kT$
Average KE $\propto T$

**10**

**(a)** molecules has component of velocity in 3 directions, hence

$c^2 = c_x^2 + c_y^2 + c_z^2$

their motions are random, so averaging gives

$< c_x^2 >=< c_y^2 >=< c_z^2 >$

Hence substitute the result of (2) into (1), then
$< c^2 >= 3 < c_x^2 >$

Hence for $N$ molecules, we have

$$pV = \frac{1}{3}Nm < c^2 >$$

**(b)** **(i)** $pV = nRT$

$$p = \frac{nRT}{L^3} = \frac{3 \times 8.31 \times (20.0 + 273.15)}{0.200^3}$$

Force on each side, $F = pA = \frac{3\times8.31\times(20.0+273.15)}{0.200^3} \times 0.200^2$

$F = 36\ 500\ N$

**(ii)** particles would attract one another so average force on wall will decrease.