1

(a) **(i)** $W = 2.60 \times 10^5 \times (3.80 - 2.30) \times 10^{-3} = 390\,J$

(ii) no change in internal energy
gas returns to its original temperature as there is no change to pressure and volume

(b)

change	q / J	w / J	ΔU / J
A to B	(+)1370	–390	(+)980
B to C	0	(+)550	(+)550
C to A	–1530	0	–1530

(c) when vapourising, greater change in separation of atoms or molecules and bonds are completely broken (compared to partially broken), hence a greater change in potential energy and internal energy.

Since there is greater change in volume, hence a greater work done.

2

(a) **(i)** Work done on gas = Area under P-V graph
$= 0.5(2.90 + 2.10) \times 10^5 \times (4.00 - 2.80) \times 10^{-4}$
$= 30.0\,J$

(ii) The increase in internal energy is the sum of the heat supplied and the work done on system.

(iii) 30.0 J. Since there is no change in temperature, there is no change in internal energy. Hence heat loss by gas is equal to work done on the gas.

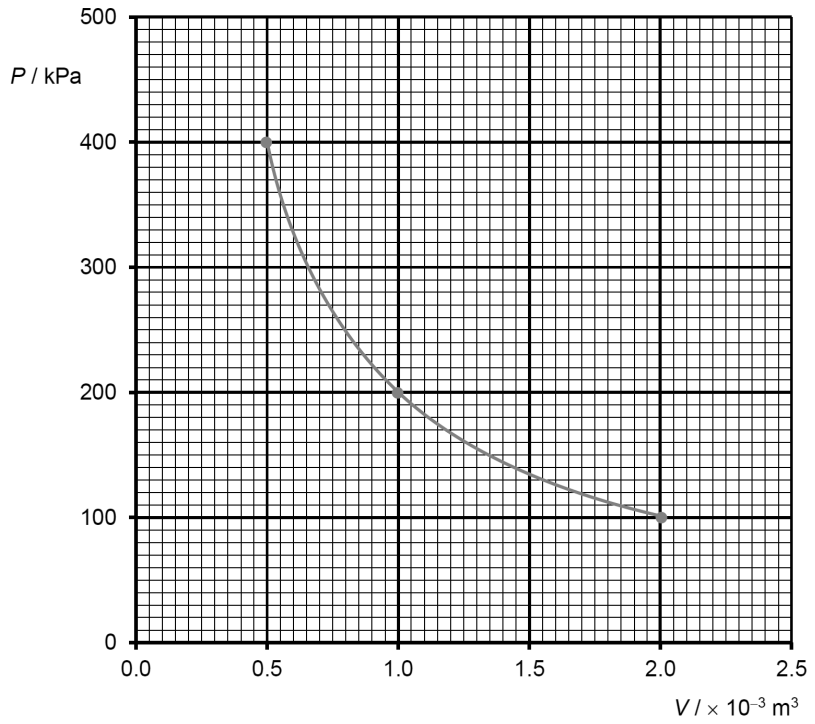
(b) **(i)** There is no heat loss to surrounding, and the work done on gas results in an increase in internal energy. This increase in internal energy is equivalent to the increase in kinetic energy, resulting in an increase in speed of the gas particles. Upon collision with the walls of the cylinder, there is a greater change in momentum. There is also higher frequency of collision between particles in the walls of the cylinder. These result in a greater force and hence greater pressure.

(ii) $\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$
$\frac{(2.10)(4.00)}{273.15 + 27} = \frac{(3.35)(3.15)}{T_f}$
$T_f = 377\,K = 104°C$

3

(a) The first law of thermodynamics states that the increase in internal energy of a system is equal to the sum of the heat supplied to the system and the work done on the system and the internal energy of a system depends only on its state.

(b) **(i)**



(ii) Since the gas expands, the area under the graph gives the work done by the gas. Since internal energy is proportional to temperature, $\Delta U = 0$ as there is no change in temperature, $0 = Q + W_{on}$ and $Q = -W_{on}$

Hence, the amount of heat supplied to the gas is equal to the area under the curve AB.

(c) For experiment 1, the volume remains constant and work done on gas = 0
Hence, increase in internal energy is equal to heat supplied. For experiment 2, the gas expands and work done on gas is negative. Hence, increase in internal energy in experiment 2 is less than that of experiment 1.

Since internal energy is proportional to temperature, increase in temperature of experiment 2 is smaller than that in experiment 1 and gas in experiment 1 will have a higher final temperature.

4

(a) **(i)** No net heat transfer between A and B hence A and B have the same temperature.

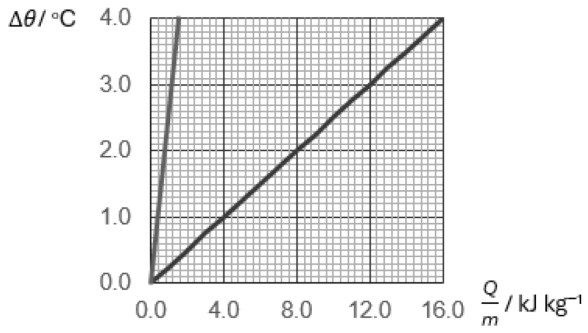
(ii) gradient of graph $= \frac{1}{c}$

$$c = \frac{1}{gradient} = \frac{1}{4/16} = 4\ kJ\ kg^{-1}\ K^{-1}$$

(iii) $m_A c_A \Delta\theta_A = m_B c_B \Delta\theta_B$ gives $c_A = \frac{1.5}{5.0} \times \frac{20}{60} \times 4000$

$$c_A = 400\ J\ kg^{-1}\ K^{-1}$$

(iv)



(b) **(i)** Δu : *increase in internal energy*
q : *heat supplied to the system*
w : *work done on the system*

(ii)

	Δu	q	w
process 1 (adiabatic)	positive	zero	positive
process 2 (constant volume)	negative	negative	zero
process 3 (isothermal)	zero	positive	negative

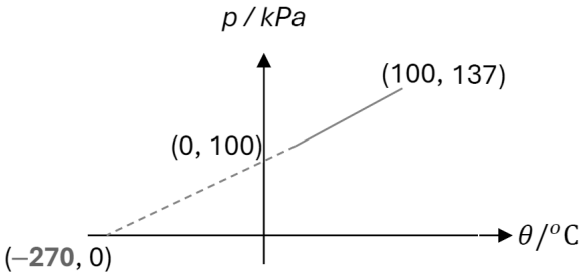
(iii) Volume remains unchanged (hence no work done)
Since heat is lost to the surrounding to melt ice-water mixture, temperature decreases and internal energy , $u \propto T$ thermodynamic temperature, so $\Delta u < 0$.

(iv) gas expands, so $w < 0$ and at constant temperature so $\Delta u = 0$

(v) for the cycle $\Delta u = 0$, so $0 = q + w$

$w = -q = -(-100 \times 334) = +3.34 \times 104\, J$

(c) **(i)**



Gradient $= \frac{0.37}{100} = 0.37$, $p = 0.37T + 100$

$when\ p = 0, T = -\dfrac{100}{0.37} = -270$

(ii) at low pressure gases are far apart and seldom interact with one another so intermolecular forces related to distance between molecules are negligible.

5

(a) **(i)** An ideal gas is one that obeys the ideal gas equation, $pV = nRT$ at all pressures, volumes and temperatures.

(ii) $pV = nRT$
$(1.00 \times 10^5)(750 \times 10^{-6}) = n(8.31)(300)$
$n = 0.030$

(b)

	work done on gas / J	heat supplied to gas / J	increase in internal energy of gas / J
Ao B	+360	0	+360
B to C	0	+670	+670
C to D	-810	0	-810
D to A	0	-220	-220

(c) the gas molecules bounce off the receding piston at lower speeds and hence lower kinetic energy. For an ideal gas, the temperature is proportional to the average kinetic energy of the molecules.

(d) The net work by the engine is positive and can be used to move the car (turn the wheels)

6

(a) **(i)** Q = mass×specific latent heat of vaporisation
$= 0.37 \times 2.3 \times 10^6$
$= 8.5 \times 10^5\ J$

(ii) $PV = nRT$
$T = (100 + 273.15) = 373.15$
$n = 0.37 \times \dfrac{1000}{18}$
$V = \dfrac{(0.37 \times 1000) \times 8.31 \times 373.15}{18 \times (1.0 \times 10^5)}$
$= 0.637\ m^3$

(iii) Work done by the water = (atmospheric pressure)(increase in volume)
$= (1.0 \times 10^5)(0.64)$
$= 6.4 \times 10^4\ J$

(iv) Work done on water is negative.
From the first law of thermodynamics,
increase in internal energy = heat supplied + work done on water
$= (8.5 - 0.64) \times 10^5$
$= 7.9 \times 10^5\ J$

(b) Kinetic energy of the molecules remains unchanged because there is no temperature change. Potential energy of the molecules increases, because molecular bonds are broken and the molecules are further apart. Hence, the internal energy of the system increases.

7

(a) **(i)** At constant pressure,
$V \propto T$
$\dfrac{V_1}{V_2} = \dfrac{T_1}{T_2}$

$\frac{500}{V_2} = \frac{250}{180}$
$V_2 = 360\ cm^3$
Work done on gas $= -p\Delta V$
$= -(80 \times 10^3)(360 - 500) \times 10^{-6}$
$= 11.2\ J$

(ii) $v = \frac{pV}{RT}$
$= \frac{(80 \times 10^3)(500 \times 10^{-6})}{8.31 \times 250} = 0.01925\ mol$

$\Delta U = \frac{3}{2}nR\Delta T = \frac{3}{2}(0.01925)(8.31)(180 - 250) = -16.8\ J$

(iii) $\Delta U = Q + W$
$Q = \Delta U - W$
$= -16.8 - 11.2$
$= -28.0\ J$

Heat loss is $28.0\ J$

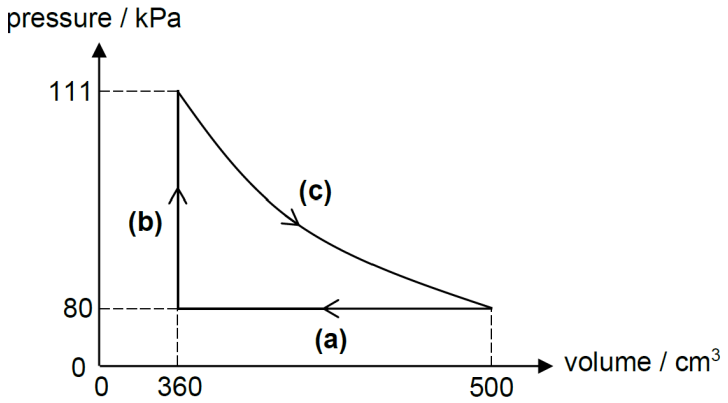
(b) **(i)** $p = \frac{nRT}{V} = \frac{(0.01925)(8.31)(250)}{360 \times 10^{-6}} = 111\ kPa$

(ii) $\Delta U = \frac{3}{2}nR\Delta T = \frac{3}{2}(0.01925)(8.31)(250 - 180) = 16.8\ J$
Since $W = 0$
$\Delta U = Q$
$Q = mc\Delta T = 16.8$
$c = \frac{16.8}{(0.23 \times 10^{-3})(250 - 180)} = 1043 \approx 1040\ J\ kg^{-1}\ K^{-1}$

(iii) $\frac{3}{2}nRT = \frac{1}{2}m_{total} < c^2 >$
$c_{r.m.s.} = \sqrt{\frac{3nRT}{m}} = \sqrt{\frac{3(0.01925)(8.31)(250)}{(0.23 \times 10^{-3})}}$
$= 722.2 \approx 720\ m\ s^{-1}$

(iv) The root-mean-square speed of the particles would remain the same. This is because the root-mean-square speed for each particle is only dependent on the temperature and not on the amount of gas, given that the mass of each particle is the same.

(c)



(d) Since the work done by the gas during expansion is greater than the work done on the gas during compression, there is a net work done by the gas in each cycle. Since there is no change in internal energy in each cycle, the gas gains heat in each cycle.

8

(a) Despite both being at $100\ °C$, steam transfers more (heat or thermal) energy to the skin than boiling water. This is because steam releases (heat) energy proportional to the specific latent heat of vaporisation during condensation in addition to the heat energy released on cooling of condensed water on skin. This is much higher than the (heat) energy released by boiling water on contact with the skin.

(b) **(i)** Using state A,

$$n = \frac{P_A V_A}{RT_A} = \frac{(20 \times 10^5 \times 1.2 \times 10^{-4})}{8.31 \times 800}$$
$$0.0361\ mol$$

(ii) 1.
$$\frac{V_A}{T_A} = \frac{V_B}{T_B}$$
$$\frac{1.28}{800} = \frac{5.2}{T_B}$$
$$T_B = 3467\ K$$

(ii) 2.
$$\Delta U = \frac{3}{2} nR\Delta T$$
$$= \frac{3}{2} \times 0.0361 \times 8.31 \times (3467 - 800)$$
$$= 1200\ J$$

(ii) 3.
$$W = -P\Delta V = -20 \times 10^5 \times (5.20 - 1.20) \times 10^{-4}$$
$$= -800\ J$$

(ii) 4.
$$\Delta U = Q + W$$
$$Q = \Delta U - W$$
$$Q = 1200 - (-800) = 2000\ J$$

(iii) From D to A, since Q is zero, $W_{DA} = \Delta U_{DA}$
$$W_{DA} = \frac{3}{2} \times 0.0361 \times 8.31 \times (800 - 226) = 258\ J$$
$$W_{net\ by\ gas} = 800 + 390 + 0 - 259 = 932\ J$$

9

(a) **(i)** Rate of mass of ice melted in Fig. 2.1 is greater due to heat from heater and surroundings compared to Fig. 2.2 from surroundings only.

(ii) Rate of ice melted = $\frac{m}{t} = gradient$

Rate of ice melted by heater = $\left(\frac{m}{t}\right)_{heater+surrounding} - \left(\frac{m}{t}\right)_{surrounding}$
$$= \frac{55 \times 10^{-3}}{4.0 \times 60} - \frac{(10 \times 10^{-3})}{4.0 \times 60}$$
$$= 1.875 \times 10^{-4}\ kg\ s^{-1}$$

$$Pt = ml_f$$
$$IV = \left(\frac{m}{t}\right) l_f$$

$l_f = \frac{(5.0)(12)}{1.875 \times 10^{-4}} = 3.2 \times 10^5\ J\ kg^{-1}$
$= 3 \times 10^5\ J\ kg^{-1}$

(b) **(i)** **1.** Speed / kinetic energy of gas molecules increase when hit by the moving piston, so temperature of gas increases

(i) **2.** The gas returned to its original temperature so no change in internal energy

(ii)

stage	thermal energy supplied to gas / J	work done on gas / J	increase in internal energy of gas / J
A to B	–702	702	0
B to C	0	844	844
C to D	936	–936	0
D to A	0	–844	–844

(iii) $\eta = \frac{net\ work\ done\ BY\ gas}{thermal\ energy\ supplied} \times 100\%$
$= \frac{(936 + 844) - (704 + 844)}{936} \times 100\%$
$= \frac{234}{936} \times 100\%$
$= 25\%$

10

(a) The internal energy of an ideal gas is just the sum of the microscopic kinetic energy, due to the random motion of its molecules since the microscopic potential energy of an ideal gas is zero.

Its internal energy depends only on its state (pressure, volume, temperature) and amount of gas.

(b) **(i)** **1.** $\Delta U = Q + W$
Since volume is constant, $W = 0$

$Q = \Delta U$
$= \frac{3}{2}p_2V - \frac{3}{2}p_1V$
$= \frac{3}{2}V(p_2 - p_1)$
$= \frac{3}{2}(2.0 \times 10^{-2})(1.5 \times 10^5 - 1.0 \times 10^5)$
$= 1500\ J$

(b) **(i)** **2.** Since $E_K = \frac{3}{2}kT_1 = 6.2 \times 10^{-21}\ J$
$T_1 = \frac{2}{3}\left(\frac{6.2 \times 10^{-21}}{1.38 \times 10^{-23}}\right) = 299.52\ K$
From $p = \frac{nR}{V}T$
$\Delta p = \frac{nR}{V}\Delta T = \frac{p_1}{V_1}\Delta T$

$$\Delta T = \frac{\Delta p}{p_1} T_1$$
$$= \frac{0.5 \times 10^5}{1.0 \times 10^5} (299.52)$$
$$= 150\ K$$

(ii) The first law of thermodynamics states that the increase in internal energy ΔU is equal to the sum of the work done on the system W and the heat supplied to the system Q, i.e. $\Delta U = Q + W$. Hence $Q = \Delta U - W$.

When a unit mass of the gas, heated under constant volume or under constant pressure, experiences a unit rise in temperature, ΔU will be the same since $\Delta U \propto \Delta T$.

When the gas is heated at constant volume, there is no work done. W is zero and $Q_V = \Delta U$.

When the gas is heated at constant pressure, work is done by the gas as it expands. W is negative and $Q_P > \Delta U$

The specific heat capacity of a gas is the heat supplied to a unit mass of the gas to cause a unit rise in its temperature. Since $Q_P > Q_V$ the specific heat capacity at constant pressure is higher than that at constant volume.