Past-paper walk-through

The first law of thermodynamics ▪ 16.2

eXercise

Questions in this set

- 9702/43 N25 Q2 ▪ 8 marks
- 9702/41 N23 Q2 ▪ 12 marks
- 9702/42 M23 Q2 ▪ 12 marks

Reading the mark scheme

- **B1** a mark that stands on its own – the point is either made or not.
- **M1** a *method* mark: the right approach, even if the number is wrong.
- **A1** an *answer* mark, and it usually needs its M mark first.
- **C1** a working mark on the way to the answer.
- // separates answers the examiner accepts equally.
- **ecf** = error carried forward: a later part is marked on your own earlier answer.
- **owtte** = "or words to that effect" – the wording need not match.

Question 2

9702/43 N25 ▪ Q2 ▪ 8 marks

(a) State **two** ways in which the first law of thermodynamics describes that the internal energy of a system may be changed.

1. 2. **[2]**

(b)(i) Use the first law of thermodynamics to explain why a bicycle pump gets hot when it is used to pump up a tyre quickly. **[3]**

(b)(ii) With reference to molecular energies, explain why the temperature of water remains at $100\ ^{\circ}\text{C}$ when it vaporises in a kettle, even though it is being heated. **[3]**

Solution 2

(a)

Mark scheme

- work done on / by system **B1**
- thermal energy supplied to / removed from system **B1**

Solution 2

(b)(i)

Mark scheme

- no thermal energy transferred to / from system (due to lack of time) B1
- work is done on the gas to compress it / to decrease its volume B1
- internal energy increases so temperature increases B1

Solution 2

(b)(ii)

Mark scheme

- (during vaporisation) molecular separation increases **B1**
- (heating causes) potential energy of molecules to increase **B1**
- kinetic energy of molecules unchanged so temperature unchanged **B1**

Question 2

9702/41 N23 ▪ Q2 ▪ 12 marks

(a) Define specific heat capacity. **[2]**

(b)(i) Show that the magnitude of the work done on the gas is 5000 J. **[1]**

(b)(ii) Explain whether the work done on the gas is positive or negative. **[2]**

(b)(iii) Determine the magnitude of the thermal energy q transferred to the gas. **[2]**

(b)(iv) Calculate the specific heat capacity of the gas for this process. Give a unit with your answer. **[2]**

(c) The gas in **(b)** is now heated at constant volume rather than at constant pressure. The increase in internal energy of the gas is the same as in **(b)**.

Use the first law of thermodynamics to explain whether the specific heat capacity of the gas for this process is less than, the same as, or greater than the answer in **(b)(iv)**. **[3]**

Solution 2

(a)

Mark scheme

- (thermal) energy per unit mass (to change temperature)
- (thermal) energy per unit change in temperature

Solution 2

(b)(i)

Worked steps

Know. When a gas changes volume at **constant** pressure, the work involved is the pressure times the volume change.
Equation. $W = p\Delta V$.
Substitute. $W = (2.0 \times 10^5) \times (0.063 - 0.038)$.
Answer. $W = 5000$ J, as required.
Check. Use the **change** in volume, not the final volume. And keep the pressure in pascals: a pressure in kPa with a volume in m^3 gives an answer a thousand times too small.

Mark scheme

- work done $= p\Delta V$
- $= (2.0 \times 10^5) \times (0.063 - 0.038) = 5000$ J

Solution 2

(b)(ii)

Mark scheme

- gas is expanding (against external pressure)
- gas does work / work is done by gas, so (work done on gas is) negative

Solution 2

(b)(iii) Worked steps

Know. The first law says the increase in internal energy equals the thermal energy supplied **plus** the work done **on** the gas: $\Delta U = q + W$.
Get the sign right. The gas **expands**, so it does work on its surroundings and the work done *on* it is negative: $W = -5000$ J.
Substitute. $7600 = q + (-5000)$.
Answer. $q = 12\,600$ J.

Solution 2

Check. More heat goes in than the internal energy rises by, because some of the energy supplied leaves again as the work of pushing the surroundings back. An answer of 2600 J means the sign of W was taken the wrong way round.

Mark scheme

- $\Delta U = q + W$
- $7600 = q + (-5000)$
- $q = 12\,600$ J

Solution 2

(b)(iv) Worked steps

Know. Specific heat capacity is the thermal energy needed per kilogram per kelvin. Use the **thermal energy** just found, not the change in internal energy.

Equation. $c = \dfrac{q}{m\Delta T}$.

Substitute. $c = \dfrac{12\,600}{0.35 \times 56}$.

Answer. $c = 640$ J kg^{-1} K^{-1}, and the unit is asked for.

Solution 2

Check. A temperature **change** is the same number in °C and in K, so no conversion is needed here. Notice this value belongs to this process only: heating the same gas at constant volume needs less energy, because then none of it is spent doing work.

Mark scheme

- specific heat capacity $= q/m\Delta T$
- $= 12600/(0.35 \times 56)$
- $= 640$ J kg^{-1} K^{-1}

Solution 2

(c)

Mark scheme

- same gain in internal energy so same temperature rise
- no change in volume so no work done
- **or**
- no work done so less thermal energy needed (for same change in internal energy)
- less thermal energy needed (for same temperature change) so lower specific heat capacity

Question 2

9702/42 M23 ▪ Q2 ▪ 12 marks

(a) State what is meant by an ideal gas. **[2]**

Question 2

9702/42 M23 ▪ Q2 ▪ 12 marks

(b)(i) Show that the helium gas behaves as an ideal gas. **[2]**

(b)(ii) The first law of thermodynamics may be expressed as

$$\Delta U = q + W$$

Use the first law of thermodynamics to explain why the temperature of the helium gas increases. **[2]**

(b)(iii) The average translational kinetic energy E_K of a molecule of an ideal gas is given by

$$E_K = \frac{3}{2}kT$$

where k is the Boltzmann constant and T is the thermodynamic temperature.

Calculate the change in the total kinetic energy of the molecules of the helium gas. **[3]**

Question 2

9702/42 M23 ▪ Q2 ▪ 12 marks

(c) The mass of nitrogen gas in another container is 24.0 g at a temperature of 27 °C. The gas is cooled to its boiling point of -196 °C. Assume all the gas condenses to a liquid.

For this change the specific heat capacity of nitrogen gas is 1.04 kJ kg^{-1} K^{-1}.

The specific latent heat of vaporisation of nitrogen is 199 kJ kg^{-1}.

Determine the thermal energy, in kJ, removed from the nitrogen gas. **[3]**

Solution 2

(a)

Mark scheme

- gas for which $pV \propto T$
- where T is thermodynamic temperature

Solution 2

(b)(i) **Worked steps**

Know. An ideal gas obeys $pV = nRT$, so for a **fixed amount** of gas the quantity $\dfrac{pV}{T}$ has the same value in every state. Show that the two states give the same number and the case is made.

Convert the temperatures. Kelvin, always: $27\ ^\circ\text{C} = 300$ K and $742\ ^\circ\text{C} = 1015$ K.

State 1. $\dfrac{(1.10 \times 10^5)(540 \times 10^{-6})}{300} = 0.198.$

Solution 2

State 2. $\dfrac{(6.70 \times 10^{6})(30 \times 10^{-6})}{1015} = 0.198.$

Answer. The two values agree, so the helium behaves as an ideal gas.

Solution 2

Check. Both temperature conversions carry a mark on their own. Using degrees Celsius makes the second value about eight times the first and would suggest, wrongly, that the gas is far from ideal.

Mark scheme

- evidence of two temperature conversions between °C and K
- two calculations shown, one for each state e.g.
- $\frac{1.10\times10^{5}\times540\times10^{-6}}{273+27} = 0.198$ and $\frac{6.70\times10^{6}\times30\times10^{-6}}{273+742} = 0.198$

Solution 2

(b)(ii)

Mark scheme

- work is done on the gas
- internal energy increases (so temperature increases)

Solution 2

(b)(iii) Worked steps

Know. Two steps. First find **how many** molecules there are, then find how much the average kinetic energy of each one changes.

Count the molecules. $pV = NkT$, so $N = \dfrac{(1.10 \times 10^5)(540 \times 10^{-6})}{(1.38 \times 10^{-23})(300)} = 1.44 \times 10^{22}$.

Energy change per molecule. $E_K = \frac{3}{2}kT$, so a temperature rise ΔT changes it by $\frac{3}{2}k\Delta T$, with $\Delta T = 1015 - 300 = 715$ K.

Multiply up. $\Delta E_K = \frac{3}{2}k\Delta T \times N = \frac{3}{2}(1.38 \times 10^{-23})(715)(1.44 \times 10^{22})$.

Answer. $\Delta E_K = 212$ J.

Check. N can be found from **either** state, since the amount of gas does not change. Using $\Delta T = 715$ is essential: putting in the two temperatures separately and subtracting the energies gives the same thing, but putting in $742 - 27$ in degrees

Solution 2

Celsius happens to give the same 715 only because a temperature *difference* is the same on both scales. **Mark scheme**

- $pV = NkT$ e.g.
- $N = \frac{1.10\times10^{5}\times540\times10^{-6}}{1.38\times10^{-23}\times300}$
- $= 1.435 \times 10^{22}$
- $\Delta E_k = (3/2)k\Delta TN$
- $= (3/2) \times 1.38 \times 10^{-23} \times (742 - 27) \times \frac{1.10\times10^{5}\times540\times10^{-6}}{1.38\times10^{-23}\times300}$
- $= 212$ J

Solution 2

(c) Worked steps

Know. Two separate processes happen, and the energy for each must be worked out and added. The gas first **cools** to its boiling point, and then **condenses** at that fixed temperature.

Convert. $m = 24.0\ \text{g} = 0.0240\ \text{kg}$, and the temperature falls by $\Delta\theta = 27 - (-196) = 223\ \text{K}$.

Cooling. $E_1 = mc\Delta\theta = 0.0240 \times 1.04 \times 223 = 5.57\ \text{kJ}$.

Condensing. $E_2 = mL = 0.0240 \times 199 = 4.78\ \text{kJ}$.

Answer. $E = E_1 + E_2 = 10.3\ \text{kJ}$.

Solution 2

Check. Forgetting the latent heat is the classic loss here, and it costs nearly half the energy. Note also that the temperature does **not** change while the gas condenses, which is why the two terms are separate and not combined into one.

Mark scheme

- $E = mc\Delta\theta$ and $E = mL$
- $\Delta\theta = (27 + 196)$ or 223
- $E = 0.0240 \times 1.04 \times (27 + 196) + 0.0240 \times 199$
- $= 10.3$ kJ