JURONG PIONEER JUNIOR COLLEGE
9749 H2 PHYSICS

FIRST LAW OF THERMODYNAMICS

Content

- **Specific heat capacity and specific latent heat**
- **Internal energy**
- **First law of thermodynamics**

Learning Outcomes

Candidates should be able to:

(a) define and use the concepts of specific heat capacity and specific latent heat.

(b) show an understanding that internal energy is determined by the state of the system and that it can be expressed as the sum of a random distribution of kinetic and potential energies associated with the molecules of a system.

(c) relate a rise in temperature of a body to an increase in its internal energy.

(d) recall and use the first law of thermodynamics expressed in terms of the increase in internal energy, the heat supplied to the system and the work done on the system.

Introduction

Heat capacity and specific heat capacity allows for the calculation of temperature changes in the interaction between two substances in thermal contact. Latent heat is used to calculate the energy required to change the phase of a substance.

Thermodynamics is the study of relationships involving heat, mechanical work, and other aspects of energy and energy transfer. The first law of thermodynamics is central to understanding thermodynamic processes which involve heat and mechanical work. This law extends the principle of conservation of energy in Newtonian mechanics, by introducing the concept of internal energy.

1 Specific Heat Capacity and Specific Latent Heat

(a) Candidates should be able to define and use the concepts of specific heat capacity and specific latent heat.

1.1 Heat capacity

Heat capacity *C* of a substance is the amount of thermal energy required to cause a unit rise in temperature of the substance.

SI units of heat capacity: $\mathbf{J\ K^{-1}}$

In symbols, we have $C = \dfrac{Q}{\Delta T}$

Hence, $\boldsymbol{Q = C\Delta T}$

where $\boldsymbol{Q}$ is the heat supplied (or removed),
$\boldsymbol{C}$ is the heat capacity,
$\boldsymbol{\Delta T}$ is the change in temperature.

Note: Heat capacity is dependent on both the mass of the body and the material of the body.

Example 1

How much heat must be supplied to an object with a heat capacity of 150 J K^{-1} in order to increase its temperature from 350 K to 400 K?

Solution:
We assume that the object does not change phase in this temperature range.

Heat supplied = $C\Delta T$ = (150)(400 – 350) = 7500 J

1.2 Specific heat capacity

The **specific heat capacity c of a substance** is defined as the thermal energy per unit mass required to raise the temperature of the substance by one unit of temperature.

SI units of specific heat capacity: **J kg^{-1} K^{-1}**

In symbols, we have $c = \dfrac{Q}{m\,\Delta T}$

Hence, $\boldsymbol{Q = mc\Delta T}$

where
$\boldsymbol{Q}$ is the heat supplied (or removed),
$\boldsymbol{m}$ is the mass of substance,
$\boldsymbol{c}$ is the specific heat capacity,
$\boldsymbol{\Delta T}$ is the change in temperature.

Note: The specific heat capacity is the heat capacity per unit mass of a substance and it is characteristic of the material of the substance.

Example 2
The heat capacity of 2.0 kg of water is 8400 J K^{-1}. What is the specific heat capacity of water?

Solution:
Specific heat capacity of water = C / m
= 8400 / 2.0 = 4200 J kg^{-1} K^{-1}

Example 3
An electrical heater supplies 1200 J of thermal energy to an aluminium cylinder of mass 32.4 g. What is the increase in temperature of the cylinder?
(Specific heat capacity of aluminium = 910 J kg^{-1} K^{-1})

Solution:
We assume that all the heat supplied by the heater goes into increasing the temperature of the aluminium cylinder.

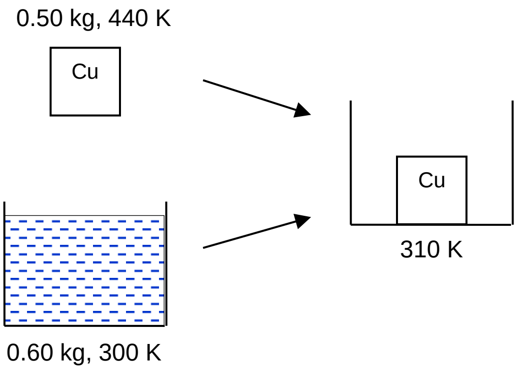
Heat absorbed by cylinder = $mc\Delta T$
$1200 = 0.0324 \times 910 \times \Delta T$
$\therefore\ \Delta T = 40.7$ K

Example 4

A block of copper of mass 0.50 kg at an initial temperature of 440 K is placed in 0.60 kg of water at 300 K. When thermal equilibrium is attained, the temperature of the system is 310 K.

How much heat is lost by the copper block and how much heat is absorbed by the water? Comment on your answers.
(Specific heat capacities: water = 4200 J kg^{-1} K^{-1}, copper = 400 J kg^{-1} K^{-1})

Solution:



Heat change of copper
= $mc\Delta T$
= 0.50 x 400 x (310 – 440)
= – 26000 J
= – 26 kJ

Since the heat change is negative, heat lost by the copper block is 26 kJ.

Heat change of water
= 0.60 x 4200 x (310 – 300)
= 25200 J = 25 kJ

Since the heat change is positive, heat gained by the water is 25 kJ.

We see that the heat lost by the copper is greater than the heat gained by the water, which means that some heat (about 1 kJ) was lost to the surroundings (i.e. the air, the container) while the copper block and water were in the process of attaining thermal equilibrium with each other.

1.3 Specific latent heat

- There are three **phases of matter**: solid, liquid and gas. Matter can change from one phase to another. It takes energy for a substance to change from solid to liquid or from liquid to gas.

- In using the concept of heat capacity, the substance changes temperature without changing its phase (i.e. it does not go from solid to liquid or from gas to solid). In using the concept of latent heat, the substance changes phase without changing temperature. The heat is termed “latent” because it is present but “hidden” as the temperature stays constant.

- Pure substances have sharp melting points and boiling points, as observed in Fig. 1.1.

- At these temperatures, when more heat is supplied to the system, the average kinetic energy of the particles does not change (**temperature remains constant**), and only the potential energy part of the internal energy of the substance increases (particles move further apart from one another).

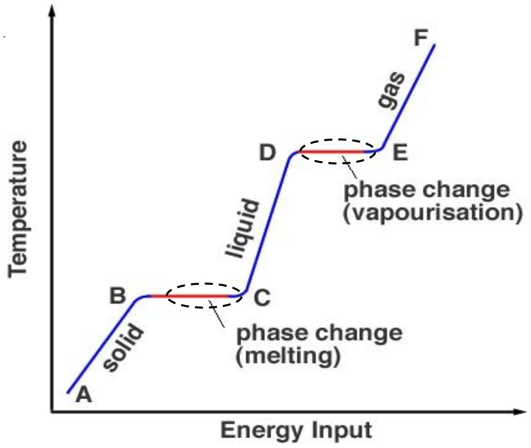


Fig. 1.1 Phase change with temperature

- At the melting point, heat supplied turns the substance from solid into liquid. At the boiling point, heat supplied turns the substance from liquid into gas. If heat is extracted from the substance, the reverse happens; there is solidification or liquefaction without a change in temperature.

> **Specific latent heat l of a substance** is the amount of thermal energy per unit mass required to change the phase of the substance at constant temperature.
>
> SI units of specific latent heat: **J kg^{-1}**
>
> In symbols, we have $Q = ml$
>
> where Q is the heat supplied (or removed) to change the phase of the substance, m is the mass of substance, l is the specific latent heat.

1.3.1 Specific latent heat of fusion

- In the process of changing from a solid into a liquid, the rigid molecular structure of the particles in fixed positions breaks down such that the particles can move around one another.

- This is known as melting, and the temperature at which this phase change occurs is the **melting point** of the substance.

- The reverse process of forming a solid from the liquid is called freezing or sometimes fusion.

> **Specific latent heat of fusion l_f** is the amount of thermal energy per unit mass required to convert a solid, at its melting point, into liquid at the same temperature, or vice-versa (from liquid to solid).

1.3.2 Specific latent heat of vaporisation

- In the process of changing from a liquid into a gas, the molecules will move from being next to one another to being very far apart.

- This process of going from liquid to gas is known as vaporisation, and the temperature at which this phase change occurs is the **boiling point** of the substance.

- The reverse process of forming a liquid from the gas is called liquefaction.

> **Specific latent heat of vaporisation l_v** is the amount of thermal energy per unit mass required to convert a liquid, at its boiling point, into gas at the same temperature, or vice-versa (from gas to liquid).

- During melting and boiling, the potential energies of the particles are increased because the forces of attraction are being overcome; particles are breaking out of their fixed positions (solid to liquid) or are dramatically increasing their separation (liquid to gas).

- Additionally, for boiling, some energy is required to do work against the atmospheric pressure because the gas occupies a much greater volume than the liquid. The gas must expand and hence there is work done **by** the gas.

Example 5

The specific latent heat of fusion of water is 330 kJ kg^{-1}. The specific heat capacity of water is 4.2 kJ kg^{-1} K^{-1}. How much heat must a refrigerator remove from 0.20 kg of water in order to turn it into ice completely, if the water was originally at (a) 0 °C ; (b) 25 °C ?

Solution:

(a) Heat removed from water at 0 °C = $m l_f$

$= (0.20)(330) = 66$ kJ

(b) Heat removed from water at 25 °C = $m l_f + m c \Delta T$

$= 66.0 + (0.20)(4.2)(25) = 87$ kJ

Example 6

The specific latent heat of vaporisation of steam is 2.26×10^6 Jkg^{-1}. The specific heat capacity of water is 4.20×10^3 J$kg^{-1}K^{-1}$. Calculate the energy required to convert 0.400 kg of water into steam, if all the water was originally at (a) 373 K ; (b) 298 K .
(Take the boiling point of water to be 373 K at standard pressure.)

Solution:

(a) If the water is already at the temperature of the boiling point,

energy required = $m l_v = (0.400)(2.26 \times 10^6) = 9.04 \times 10^5$ J

(b) For this case, energy is needed to increase the temperature of the water to its boiling point before boiling can occur.

Total energy required = $m l_v + m c \Delta T$

$= 9.04 \times 10^5 \text{ J} + (0.400)(4.20 \times 10^3)(373 - 298)$

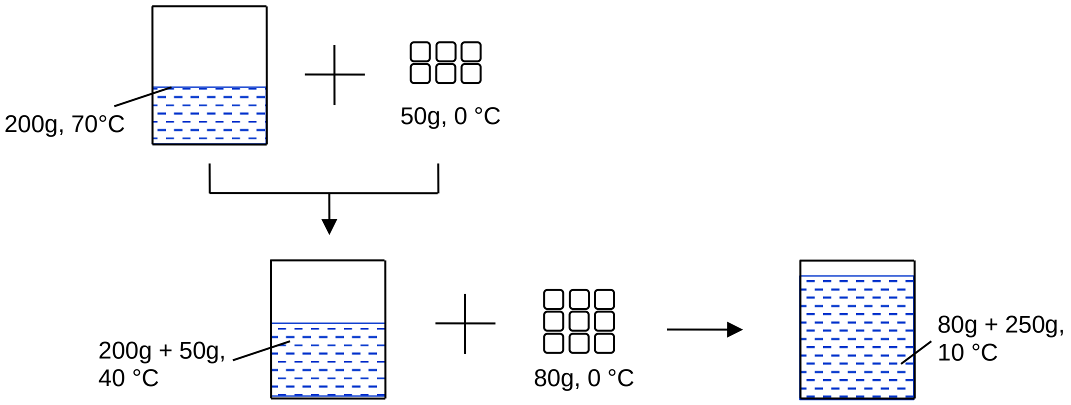
$= 1.03 \times 10^6$ J

Example 7

50 g of ice at 0 °C is added to 200 g of water initially at 70 °C in a vacuum flask. When all the ice has melted, the temperature of the flask and its contents is 40 °C. When a further 80 g of ice at 0 °C has been added and has all melted, the temperature of the whole system becomes 10 °C.

Calculate the specific latent heat of fusion of ice neglecting any heat loss to the surroundings. (Specific heat capacity of water is 4200 J $\text{kg}^{-1}\ \text{K}^{-1}$.)

Solution:
For this question, we must not neglect the heat capacity of the container. The temperature of the container and contents happen to drop by 30 °C each time ice is added, so the heat absorbed by the container ($Q = C\Delta T$) is the same each time. We shall let this heat absorbed be H.



For both processes, since we assume no heat loss to the surroundings, we use the equation

Heat gained by ice = Heat lost by water and container

$$m_i l_f + m_i c_w \Delta T_i = m_w c_w \Delta T_w + H$$

$$(0.050)\, l_f + (0.050)(4200)(40 - 0) = (0.200)(4200)(70 - 40) + H \quad \text{----------} \quad (1)$$

$$(0.080)\, l_f + (0.080)(4200)(10 - 0) = (0.250)(4200)(40 - 10) + H \quad \text{----------} \quad (2)$$

Equation (2) – (1): $l_f = 378\ \text{kJ kg}^{-1}$

2 Internal Energy

(b) Candidates should be able to show an understanding that internal energy is determined by the state of the system and that it can be expressed as the sum of a random distribution of kinetic and potential energies associated with the molecules of a system.

(c) Candidates should be able to relate a rise in temperature of a body to an increase in its internal energy.

2.1 Internal energy

The internal energy *U* of a system is the sum of a random distribution of kinetic and potential energies associated with the molecules of a system**.** It is determined by the state of the system.

$$U = \sum K.E._{microscopic} + \sum P.E._{microscopic}$$

2.1.1 Microscopic kinetic energy

- The microscopic KE arises from the continuous random motion of the particles.

- For monatomic gases, the microscopic KE refers to translational motion of atoms.

- For diatomic and polyatomic gases, the kinetic energies include other forms such as rotational and vibrational kinetic energies of the molecules.

The **temperature** T of a gas is **a measure of the mean kinetic energy** of particles in the gas.

2.1.2 Microscopic potential energy

- The microscopic PE arises from the intermolecular forces and it depends on the type of bond between the particles and the intermolecular separation. The potential energy of the molecules is due to the **work done to move them further apart against this *intermolecular attraction*.**

- Hence, potential energy is the highest when the particles are separated far apart, and the potential energy drops as the particles are nearer. This is similar to gravitational potential energy: the potential energy is high when the object is high above the ground and low when the object is subjected to the gravitational attractive force and moves nearer to the ground.

- Potential energy is taken to be zero (highest) when the particles are separated far apart (intermolecular forces are so small they are negligible like in an ideal gas).

- This means that the potential energy is negative (smaller than zero) when the particles are nearer to one another.

2.2 Melting and boiling

- The microscopic KE and microscopic PE of particles can be used to explain why melting and boiling take place without a change in temperature.

Melting

The heat that flows into the melting solid increases only the **potential energy** component of the internal energy, and not the **kinetic energy** of the particles. Thus the temperature does not change and the heat required for melting is known as the latent heat of fusion.

Boiling

The heat that flows into the boiling liquid increases only the **potential energy** component of the internal energy, and not the **kinetic energy** of the particles. Thus the temperature does not change and the heat required for boiling is known as the latent heat of vaporisation.

2.3 Internal energy of an ideal gas

- **For an ideal gas, the microscopic potential energy is taken to be zero** because the forces of attraction between the particles are assumed to be *zero*. The total kinetic energy of the molecules in an ideal gas as they move around the container is thus the **internal energy *U*** of the gas.

- For a monoatomic ideal gas of N atoms,

  internal energy U = total microscopic KE = $N\langle KE\rangle = N\left(\frac{3}{2}kT\right) = \frac{3}{2}NkT$

  Therefore, $\boxed{U = \frac{3}{2}NkT = \frac{3}{2}nRT}$ for a **monatomic ideal gas**.

- Using the equation of state for an ideal gas, the internal energy can also be expressed as

  $U = \frac{3}{2}pV$.

3 First Law Of Thermodynamics

3.1 First law of thermodynamics

(d) Candidates should be able to recall and use the first law of thermodynamics expressed in terms of the increase in internal energy, the heat supplied to the system and the work done on the system.

- The first law of thermodynamics is a version of the **law of conservation of energy**.

- If energy flows from outside the system into the system, then the system must have increased its energy by that same amount. Quite simply, if the **internal energy** of the system changes, it is due to the combination of the net heat supplied to the system (from outside) and the net work done ON the system (from outside).

- The first law of thermodynamics states that the **<u>increase</u>** in internal energy of a system is the sum of the heat **<u>supplied to</u>** the system and the **<u>work done on</u>** the system.

increase in internal energy of system = heat **supplied** to system + work done **ON** the system

$$\Delta U = Q + W$$

The first law applies to all systems, whether solid, liquid or gas. In essence, it says that the energy that is put **inside** the system must come from **outside**.

	positive	**negative**
ΔU	**increase** in internal energy	**decrease** in internal energy
Q	heat **supplied** to system	heat **lost** by system
W	work done **on** system	work done **by** system

3.2 Work done ON a system

- The work done on a system is defined to be the **work done by the external forces exerted on the system**.

- The work done BY the system is the negative of the work done ON the system. If 1 J of work is done ON the system, we say that −1 J of work is done BY the system since the system did not actually do the work (work was done on the system instead).

Work done is related to the **change in volume** of the system. The SI unit for work done is the joule (J).

Work done ON the system is **positive** when the volume of the system decreases (compression). When the volume of the system increases (expansion), the work done BY the system is **positive**.

Work done ON a system, $W = -p\Delta V$

where p is the pressure of the system, and ΔV is the change in volume of the system (final volume minus initial volume). ΔV is positive for an expansion, in which case W is negative.

- W is more commonly calculated for gases, which expand and contract to a much higher degree.
- Let us consider a simple system to identify the key ideas behind the formula. This derivation follows directly from the definition of work done by a force.

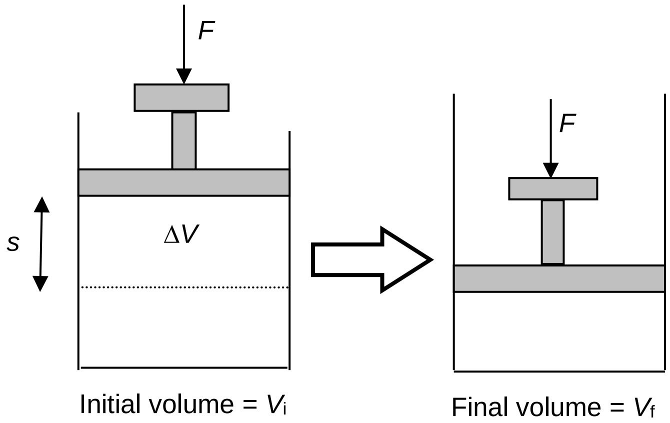


Fig. 3.1 Work done on system

- Fig. 3.1 shows a closed container filled with a fixed amount of gas. The piston is assumed to form a perfect seal. The volume of the gas changes from V_i to V_f when the piston is lowered. The gas is being compressed.
- The volume of the gas is reduced but the **pressure is kept constant.** This can be achieved by making the distance travelled by the piston, Δs, very small. This way the increase in pressure is small and can be assumed to be negligible.

Graphically, we can represent the process by plotting pressure (p) against volume (V). These are known as 'p-V graphs' (See Fig. 3.2).

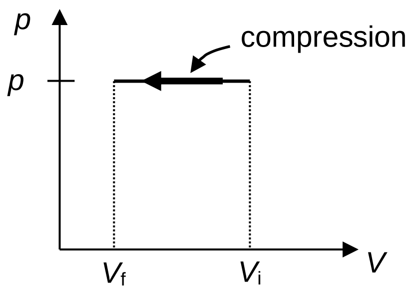


Fig. 3.2 p-V graph of compression at constant pressure

- The force F moves through a distance s, and in this example F is constant (Fig. 3.1). Let the cross-sectional area of the container be A. The pressure of the gas is given by $p = \frac{F}{A}$ so the work done by the force is given by:

$$\begin{aligned}\text{work done} &= Fs \\ &= (pA)s \\ &= p(As) \\ &= p\,|\Delta V| = -p\Delta V\end{aligned}$$

- We used the modulus symbol to take the absolute value of ΔV. Since $\Delta V = V_f - V_i$ is negative for compression, the work done on the system is $W = p\,(-\Delta V) = -p\,\Delta V$ as earlier stated without proof. Importantly, note that the work done is positive in this case even though there is a negative sign in the formula.

- The same formula applies for a gas expanding at constant pressure (See Fig. 3.3).

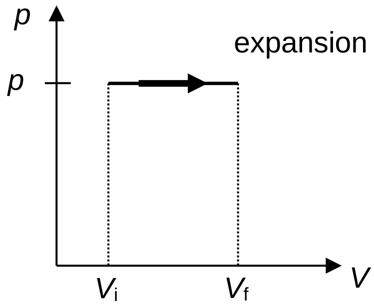


Fig. 3.3 *p*-*V* graph of expansion at constant pressure

The work done by the force on the piston is negative because the piston moves up even though the force exerted is downwards. $\Delta V = V_f - V_i$ is positive for an expansion, and we have that work done ON the system, $W = -p\,\Delta V$ is negative.

This table summarises the signs involved:

	compression	**expansion**
work done **ON** system	+	–
work done **BY** system	–	+

3.3 Work done from *p*-*V* graphs

- The formula above only applies when the pressure is kept constant.

- In general, work done on the system is found by taking the area under the pressure-volume graph (See Fig. 3.4).

- The simplest way to figure out whether work done on the system is positive or negative is to see if the volume decreases or increases respectively.

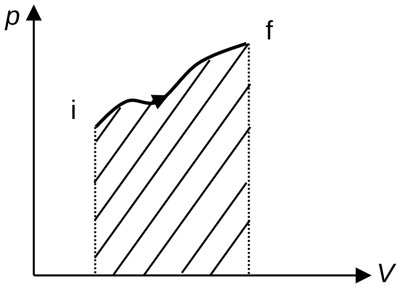


Fig. 3.4 Area under *p*-*V* graph

<u>Legend</u>
'i' represents the initial state of the system.
'f' represents the final state of the system.
The arrow shows the direction of change.

Work done ON system
= Area between the *p*-*V* graph and the volume axis

Work done ON the system is '+ve' for compression and '–ve' for expansion.

3.4 *p-V* graph and isotherms

- From the ideal gas equation, if T is constant, then pV is constant.
- This means that for a gas at constant temperature, $p \propto \frac{1}{V}$.
- Examples of constant temperature curves (*isothermal* lines, or isotherms) are shown in Fig. 3.5. The curves that are at higher temperature are "higher up", further from the axes.

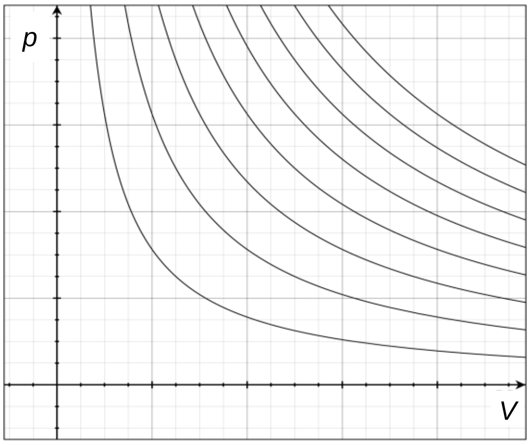
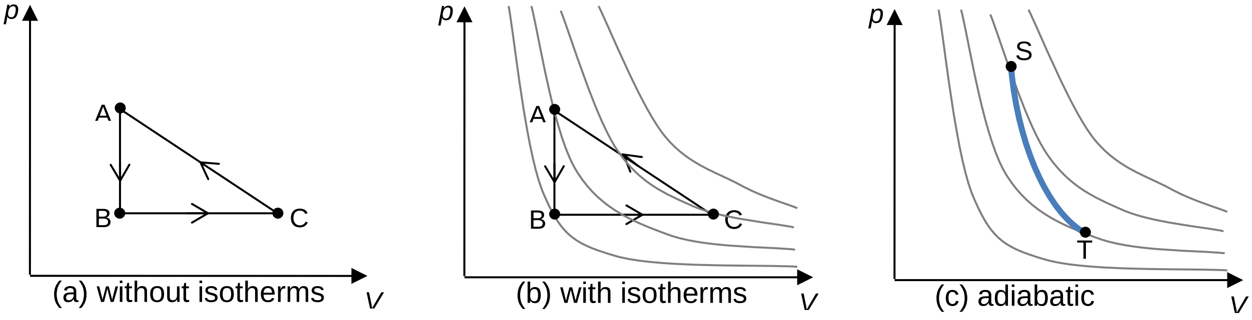


Fig. 3.5 Isothermal lines on *p*-*V* graph

- Points along the same isotherm show different states of the gas all having the same temperature. Points on different isotherms show states at different temperatures. Isotherms do not intersect.

3.4.1 Other processes on the *p-V* graph

- Fig. 3.6(a) shows a gas at an initial state A undergoing a drop in pressure to state B. The volume remains constant throughout the process and the line AB is called an *isochoric* (constant volume) line. With the isotherms added as shown in Fig. 3.6(b), it is obvious that the temperature decreases during this isochoric process.
- The process from state B to C is an expansion at constant pressure. The line BC is an *isobaric* (constant pressure) line and the temperature increases during this process.

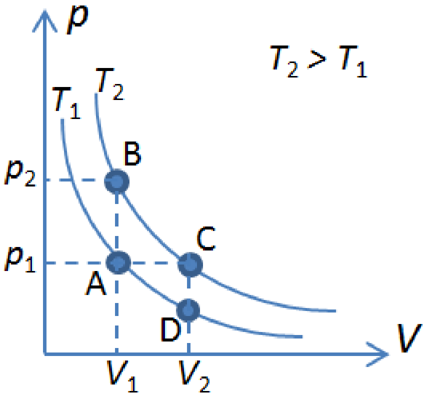


(a) without isotherms (b) with isotherms (c) adiabatic

Fig. 3.6 Other processes on a *p*-*V* graph

- The final process from state C to A is a contraction with an increase in pressure and a drop in temperature. The process A → B → C → A is called a *cyclic* process as it returns the gas to its original state A.
- It is advisable not to assume all curved lines on a *p*-*V* graph are isotherms unless it **is explicitly stated**. For e.g., an *adiabatic* process is a process where no heat is lost or gained by the gas (eg: gas kept in a thermally insulated container). The line ST shown in Fig. 3.6(c) shows an adiabatic curve which is steeper than an isotherm.

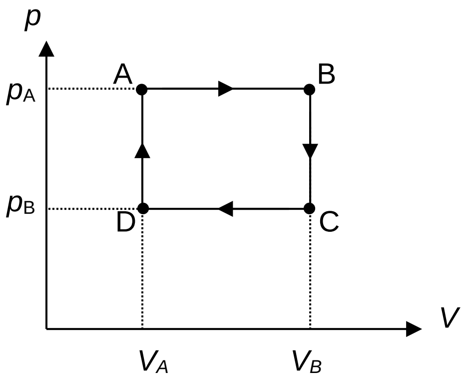
3.4.2 Summary of processes represented by *p-V* graph



process	meaning	result
Isothermal B→C & A→D: isothermal expansion. C→B & D→A: isothermal contraction.	The temperature of the system remains constant. The change in state occurs along an isotherm. Constant temperature implies constant mean kinetic energy and hence constant total kinetic energy of the molecules for a fixed amount of gas. For an ideal gas, it means constant internal energy. Both the pressure p and volume V change, but the product pV is constant, that is, $p_2V_1 = p_1V_2$.	$\Delta U = 0$
Isobaric A→C: isobaric expansion. C→A: isobaric contraction.	The pressure of the system remains constant. Constant pressure means the system's environment is not changing e.g. the system experiences atmospheric pressure. The final state is on a different isotherm, meaning the temperature changes. The volume also changes.	$W = -p\Delta V$
Isochoric / Isovolumetric A→B: isochoric heating. B→A: isochoric cooling.	The volume of the system remains constant. Constant volume usually means the system is confined in a rigid container, implying zero work done. The final state is on a different isotherm, meaning the temperature changes. The pressure also changes.	$W = 0$
Adiabatic B→D: adiabatic expansion. D→B: adiabatic contraction.	There is no heat exchange between the system and its surroundings, that is, no heat is supplied to or removed from the system. During an adiabatic expansion (B→D), the system does work by pushing back its surroundings. Since no heat is supplied to the system for it to do the work, it uses its internal energy, thus temperature decreases. The final state D is at a lower isotherm than the initial state B. The pressure also decreases. The paths for adiabatic changes are steeper than isotherms.	$Q = 0$
Cyclic A→B→C→D→A	The process returns the system to the original state (same pressure, same volume and same temperature). There will be no change in internal energy. Work done on the system = Area enclosed by the p-V graph	$\Delta U = 0$ $W \neq 0$ Work = Area enclosed by the p-V graph

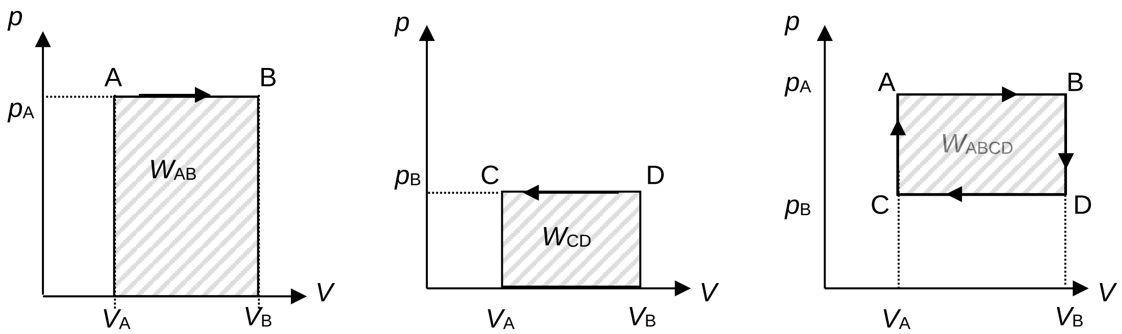
Example 8

For the cyclic process below, show the areas that represent the work done on the gas during parts AB and CD on a *p*-*V* graph.



Solution:

A **cyclic process** is a process that starts and ends at the same point in the graph (A → B → C → D → A). At the end of the process, the system is at the same state (*p*, *V*, *T*) it started with.



(a) (b) (c)

There is no work done during the stages B → C and D → A because the volume is constant during those stages.

(a) For the process A → B, work done ON the gas = $-|W_{AB}|$

(b) For the process C → D, work done ON the gas = $+|W_{CD}|$

(c) Net work done ON the gas, $W_{ABCD} = -|W_{AB}| + |W_{CD}|$

Note that the magnitude of the net work done on/by the gas in any closed cycle is equal to the area enclosed by the cycle in a p-V diagram. You still need to decide the overall sign.

(If the question asked for the work done **BY** the gas, the answers would all be of opposite sign.)

Example 9

An ideal gas of volume 1.5×10^{-3} m^3 and at a pressure of 1.0×10^5 Pa is supplied with 70 J of thermal energy. Its volume increases to 1.7×10^{-3} m^3 with the pressure remaining constant. What is the increase in the internal energy of gas?

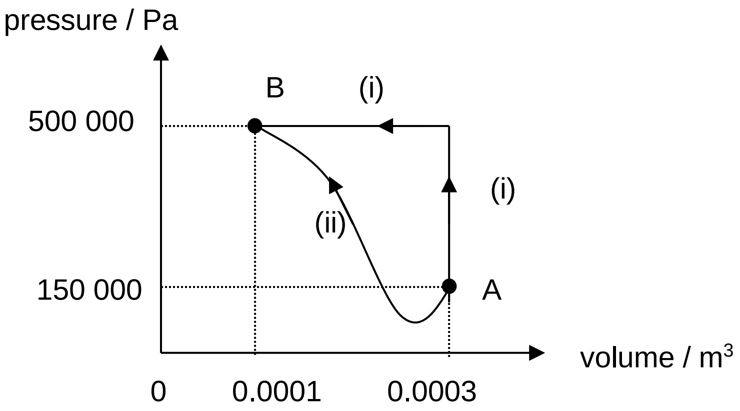
Solution:
Since the volume is increased, the work done ON the gas is negative. We must check that that is the case in our answer.

Work done on the gas, $W = -p\,\Delta V$
$= -(1.0\times10^5)(1.7\times10^{-3} - 1.5\times10^{-3})$
$= -20$ J

By the first law of thermodynamics, the increase in internal energy is given by
$\Delta U = Q + W$
$= +70 + (-20)$
$= +50$ J

Example 10

A gas can have its state changed from A to B by two different paths, as shown below.



(a) If path (i) requires the input of 60 J of heat, determine
  (i) the work done on the gas in changing its state along path (i);
  (ii) the increase in internal energy between states A and B.
(b) What can be deduced about the work done on the gas, the heat supplied to the gas and the increase in the internal energy of the gas if path (ii) is used?

Solution:
(a) (i) For path (i), there is positive work done on the gas at a pressure of 500 000 Pa.
$W = -p\,\Delta V$
$= -(500\,000)(0.0001 - 0.0003)$
$= +100$ J

(ii) We are given that for path (i), $Q = +60$ J
$\therefore \Delta U = Q + W$
$= +60 + 100 = +160$ J

(b) The initial and final states are the same for both path (i) and (ii). Since internal energy depends only on the **state** and not the process or method of getting from one state to another, state A has a certain initial internal energy and state B has a certain final internal energy. The initial internal energy for paths (i) and (ii) are the same, and likewise for the final internal energy. Hence from (a), $\Delta U_{AB} = +160$ J if path (ii) is used.

Area under path (ii) is smaller than the area under path (i) on the p-V graph.
Therefore, work done on the gas is smaller in path (ii), so from (a), $W < 100$ J.

We deduce the heat supplied from what we know about the other 2 terms in the first law. Using $\Delta U = Q + W$, we see that if ΔU is the same but W is smaller, then the heat supplied to the gas, Q, must be larger for path (ii) than for path (i), i.e. $Q > 60$ J.

References

1 Pacific Physics A-level Vol. 1 (Poh Liong Yong), pg 403 – 430.
2 A-level Physics (Roger Muncaster), pg 228 – 348.
3 Advanced Physics (Steve Adams, Jonathan Allday), pg 286 – 305.
4 Principles of Physics (Frederick J. Bueche), pg 299 – 316.
5 Physics (Robert Hutchings), pg 446 – 455.
6 Fundamentals of Physics (David Halliday, Robert Resnick, Jearl Walker), 7th Edition Extended, pg 476 – 497.