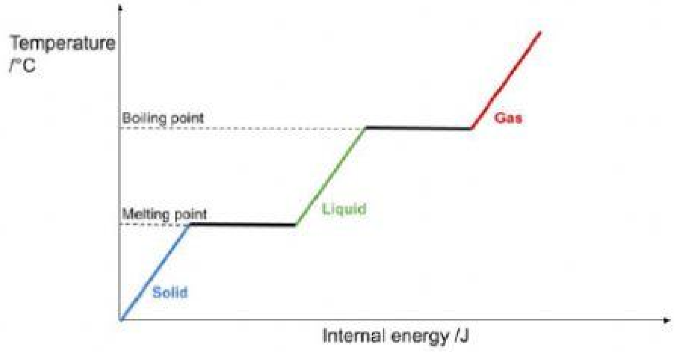
INTERNAL ENERGY & FIRST LAW OF THERMODYNAMICS

The **internal energy** of a body is equal to the **sum of all of the kinetic energies and potential energies of all its particles**. The kinetic and potential energies of a body are **randomly distributed**.

When the **state of a substance is changed, its internal energy also changes**. This is because the **potential energy of the system changes, whilst the kinetic energy of the system is kept constant.** This can be demonstrated by measuring the temperature of water as it boils and melts:

- **Boiling -**
  The temperature increases up until 100°C, after which the energy gained through heating the water is no longer used to increase the temperature (and therefore kinetic energy), but instead is **used to break bonds between water molecules** so it can change state to water vapour, and **so the potential energy is increased**.
- **Melting -**
  The temperature of ice increases up until 0°C, after which the energy gained through heating the water is no longer used to increase the temperature, but instead is **used to break bonds between water molecules** so it can change state to liquid water, and **so the potential energy is increased.**

Below is a graph showing how the internal energy of a substance varies with temperature:



The **internal energy of a body is dependent on its state**, since the state of a body is determined by the value of the kinetic energies of particles.

The **internal energy** of a system can be **increased** in two ways:

1. **Do work** on the system to transfer energy to it (e.g moving its particles/changing its shape)
2. **Increase the temperature** of the system

The **kinetic energy of particles is directly proportional to their temperature**, therefore an increase in temperature means an increase in the average kinetic energy and so an increase in internal energy.

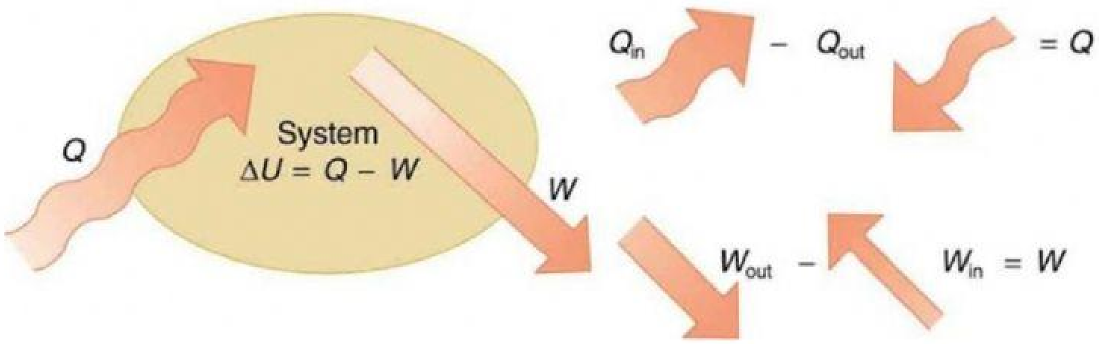
The **first law of thermodynamics** describes the conservation of energy in a system where energy can be transferred through doing work or heating. It states that the **increase in internal energy** of a system **is equal to the sum** of the **energy transferred to it through heating and work done on the system**. It is given by the following equation:

$$\Delta U = Q + W$$

In the above equation:

- **ΔU** represents the **increase** in internal energy, if **ΔU is negative**, the internal energy will **decrease**.
- **Q** is the energy transferred **to** the system through heating, therefore if **Q is negative**, energy is transferred **away** from the system through cooling.
- **W** is the work done **on** the system (occurs when a gas is **compressed**), therefore if **W is negative**, work is done **by** the system (occurs when a gas **expands**).

The diagram below shows a system which **does work**. This is shown by the fact that W is negative.



The specific latent heat of vaporisation is usually **larger** than the specific latent heat of fusion of the same substance as **more intermolecular bonds must be broken (so more work must be done)** to change a substance from a liquid to a gas, than when changing from a solid to a liquid.