Module 5 Section 1: Thermal Physics

Specific Heat Capacity

The **specific heat capacity** (c) of a substance is the amount of energy needed to raise the temperature of 1 kg of the substance by 1 K (or 1°C).

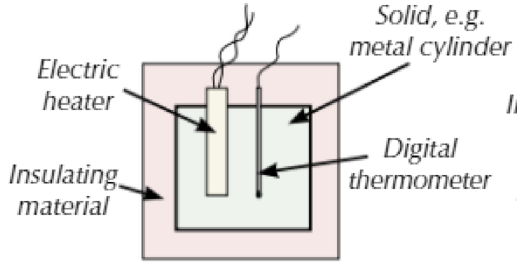
Or put another way:

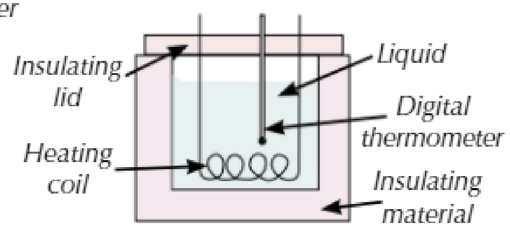
energy change = mass × specific heat capacity × change in temperature

$$E = mc\Delta\theta$$

- E = energy change in J
- c = specific heat capacity in $J\,kg^{-1}\,K^{-1}$ or $J\,kg^{-1}\,°C^{-1}$
- m = mass of substance in kg
- $\Delta\theta$ = temperature change in K or °C

Measuring Specific Heat Capacity





Specific Latent Heat

The **specific latent heat** (L) of fusion or vaporisation is the quantity of thermal energy required to change the state of 1 kg of a substance.

Which gives:

energy change = mass × specific latent heat

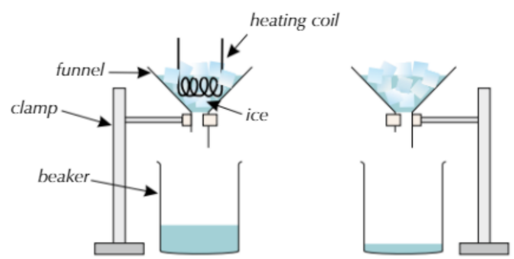
Or in symbols:

$$E = mL$$

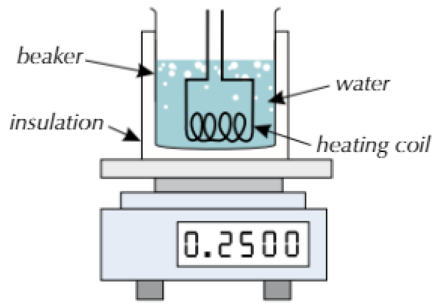
- E = energy change in J
- L = specific latent heat in $J\,kg^{-1}$
- m = mass of substance in kg

Measuring Specific Latent Heat of fusion

Fusion



Vaporisation



Gas Laws—Boyle's Law

At a constant temperature the pressure p and volume V of an ideal gas are inversely proportional.

$$pV = \text{constant}$$

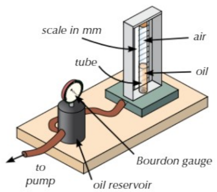


Figure 3: Experimental set-up for investigating Boyle's Law.

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Internal Energy

Internal energy:

The internal energy, U, is the sum of the kinetic and potential energies of the molecules of a system.

For a monatomic gas, there is no potential energy, so the internal energy is equal to the total translational kinetic energy of the gas (see component 3.3).

$$U = \frac{3}{2}nRT = \frac{3}{2}NkT$$

where n is the number of moles of the gas and N is the total number of molecules of gas.

Absolute zero:

Absolute zero is the temperature where the internal energy of a system **is at a minimum**.

For a gas, this is the temperature at which $pV = 0$. Therefore, it has minimum kinetic energy (and internal energy).

Heat flow:

Heat energy is always transferred from a region of **higher temperature to lower temperature**. It is unreasonable to describe the heat as being in a system as heat will always **enter or leave a system through its boundary or wall** depending on whether the temperature on the other side of the boundary is higher or lower.

If the temperature is the same on both sides of the boundary it is said to be in **thermal equilibrium** and there is no heat flow between them.

The mole

The mole:

The mole is the S.I. unit of an 'amount of substance'. It is the amount containing as many particles (e.g. molecules) as there are atoms in 12 g of carbon-12.

In 1 mole of water there are 6.02×10^{23} water molecules, this is the same number as of copper atoms in 1 mole of copper. It is known as the **Avogadro constant** and it represents **the number of particles in 1 mole of a substance**.

$$N_A = 6.02\times10^{23}\ \text{mol}^{-1}$$

Ideal Gas

Ideal gases:

Experiments investigating gas' behaviour discovered:

- Pressure $(p) \propto \frac{1}{\text{Volume }(V)}$ at a constant temperature (T)
- Pressure ∝ Temperature at a constant volume
- Volume ∝ Temperature at a constant pressure

These relationships can be combined into 1 equation:

$$pV = nRT$$

where n is the number of moles of gas and R is the molar gas constant = 8.31 J mol^{-1} K^{-1}.

Alternatively, it can be written as:

$$pV = NkT$$

where N is the number of particles and k is the Boltzmann constant = 1.38×10^{-23} J K^{-1}.

These equations are for ideal gases, but a very close approximations for the behaviour of real gases.

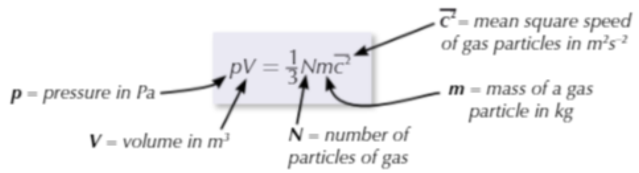
Kinetic Theory

In an ideal gas it is assumed that the molecules **collide with no loss of kinetic energy**, they only exert **forces on each other during collisions** and that the molecules are so small it can be assumed that they **take up no space**.

These assumptions also apply to the kinetic theory of gases:

- The time taken for a collision is negligible.
- The molecules move with constant velocity between collisions.
- The number of molecules is large, with a large number of collisions.
- The motion of the molecules is evenly distributed in all directions.
- **There is a random distribution of energy among the particles.**

Pressure



Internal energy of an ideal gas

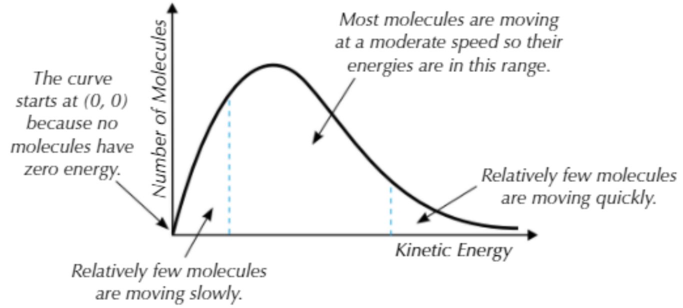


Figure 1: *The Maxwell-Boltzmann distribution of kinetic energies in a gas.*