

# University Physics with Modern Physics - Modern Physics by Young and Freedman Notes

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## 17 Temperature and Heat

### 17.1 Temperature and Thermal Equilibrium

- The **zeroth law of thermodynamics** states: If  $C$  is initially in thermal equilibrium with both  $A$  and  $B$ , then  $A$  and  $B$  are also in thermal equilibrium with each other.
- Two systems are in thermal equilibrium iff they have the same temperature.

### 17.2 Thermometers and Temperature Scales

- Water freezes at  $0^{\circ}\text{C}$  or  $32^{\circ}\text{F}$  and boils at  $100^{\circ}\text{C}$  or  $212^{\circ}\text{F}$ .
- A temperature measurement is denoted  $x^{\circ}\text{C}$  (“ $x$  degrees Celsius”) whereas a temperature interval is denoted  $x^{\circ}\text{C}^{\circ}$  (“ $x$  Celsius degrees”).

### 17.3 Gas Thermometers and the Kelvin Scale

- Under the **Kelvin** temperature scale temperature differences are equal to those of the degrees Celsius scale, but the zero is equal to  $-273.15^{\circ}\text{C}$ . This is known as **absolute zero** where molecules have their lowest possible kinetic and potential energies.
- The ratio of two temperatures in the Kelvin scale equals the ratio of the corresponding pressures in a constant-volume gas thermometer

$$\frac{T_2}{T_1} = \frac{p_2}{p_1}.$$

### 17.4 Thermal Expansion

- Materials expand when their temperatures increase.
- Expansion in a single dimension is described by the equation

$$\Delta L = \alpha L_0 \Delta T$$

where  $\Delta L$  is the change in length,  $\alpha$  is the **coefficient of linear expansion**,  $L_0$  is the original length, and  $\Delta T$  is the change in temperature.

- Expansion in three dimensions (volume expansion) is described by the equation

$$\Delta V = \beta V_0 \Delta T$$

where  $\Delta V$  is the change in volume,  $\beta$  is the **coefficient of volume expansion** (equal to  $3\alpha$ ),  $V_0$  is the original volume, and  $\Delta T$  is the change in temperature.

- If the ends of a material are fixed in place, changes in temperature can induce **thermal stresses** that can damage the material. The magnitude of these stresses is given by

$$\frac{F}{A} = -Y\alpha\Delta T.$$

## 17.5 Quantity of Heat

- Energy transferred as a result of a temperature difference is called **heat**.
- The **specific heat** of a material is the amount of energy required to raise the temperature of one unit of mass of the material by one unit of temperature, e.g. 1 kg by 1 K. It has units like J/(kg K).
- The specific heat of water is

$$4190 \text{ J}/(\text{kg K}) \text{ or } 1 \text{ cal}/(\text{g } ^\circ\text{C}).$$

- The energy required to change the temperature of a material is given by

$$Q = mc\Delta T$$

where  $m$  is the mass of the material,  $c$  is its specific heat, and  $\Delta T$  is the change in temperature.

- The **molar mass** of a substance is the mass of one mole.
- The total mass of a material  $m$  is equal to the mass per mole  $M$  times the number of moles  $n$

$$m = nM.$$

- The energy required to change the temperature of a certain number of moles of a substance is

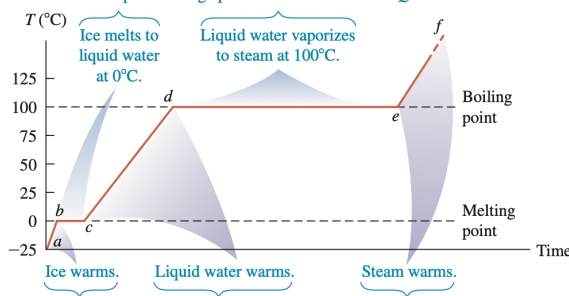
$$Q = nC\Delta T$$

where  $C = Mc$  is the **molar heat capacity**.

## 17.6 Calorimetry and Phase Changes

- A **phase** is a specific state of matter, e.g. solid, liquid, or gas.
- A **phase change** or **phase transition** is a transition from one phase to another.
- For a given pressure, phase change takes places at a definite temperature.
- While a substance is undergoing a phase change, any added or removed energy will affect the progress of the phase change but won't change the temperature.

**Phase of water changes.** During these periods, temperature stays constant and the phase change proceeds as heat is added:  $Q = +mL$ .



**Temperature of water changes.** During these periods, temperature rises as heat is added:  $Q = mc\Delta T$ .

$a \rightarrow b$ : Ice initially at  $-25^\circ\text{C}$  is warmed to  $0^\circ\text{C}$ .  
 $b \rightarrow c$ : Temperature remains at  $0^\circ\text{C}$  until all ice melts.  
 $c \rightarrow d$ : Water is warmed from  $0^\circ\text{C}$  to  $100^\circ\text{C}$ .  
 $d \rightarrow e$ : Temperature remains at  $100^\circ\text{C}$  until all water vaporizes.  
 $e \rightarrow f$ : Steam is warmed to temperatures above  $100^\circ\text{C}$ .

- The heat transfer required for a material to undergo a phase change is given by

$$Q = \pm mL$$

where the  $\pm$  indicates that heat may need to be added or removed depending on the direction of the phase change (e.g. energy must be added to melt ice),  $m$  is the mass of the material, and  $L$  is the latent heat associated with the phase change.

- When a material is freezing or melting,  $L = L_f$  the **latent heat of fusion**.
- When a material is condensing or vaporising,  $L = L_v$  the **latent heat of vaporisation**.
- When a material sublimates (changes directly from a solid to a gas, skipping liquid) or deposits/desublimates (changes directly from a gas to a solid, skipping liquid),  $L = L_s$  the **latent heat of sublimation**.
- When a material burns,  $L = L_c$  the **latent heat of combustion**.
- For any given material at any given pressure, the freezing temperature is the same as the melting temperature. This is called **phase equilibrium**. Similarly the condensing temperature is the same as the vaporisation temperature.

## 17.7 Mechanisms of Heat Transfer

- **Conduction** is a mechanism of heat transfer where the molecules in an area of high temperature have greater kinetic energy, they bump neighboring molecules which increases their kinetic energy, and so on spreading the heat through the material.
- The direction of heat flow is always from higher to lower temperature.
- When a quantity of heat  $dQ$  is transferred through a material in time  $dt$  we say the rate of heat flow or the **heat current** is

$$H = \frac{dQ}{dt}.$$

- If a rod has cross sectional area  $A$ , length  $L$ , one end is held at temperature  $T_H$ , and the other is held at  $T_C$  where  $T_H > T_C$ , the heat current is

$$H = \frac{dQ}{dT} = kA \frac{T_H - T_C}{L}$$

where  $k$  is the **thermal conductivity** of the material and  $(T_H - T_C)/L$  is the temperature difference per unit length or the magnitude of the **temperature gradient**.

- **Convection** is the transfer of heat by mass motion of fluid from one region of space to another, e.g. ducted cooling/heating. If the fluid is circulated by a blower or a pump the process is called **forced convection**; if the flow is caused by differences in density due to thermal expansion, such as hot air rising, the process is called **free convection**.
- **Radiation** is the transfer of heat by electromagnetic waves such as visible light, infrared, and ultraviolet radiation.
- The wavelength of the radiation depends on temperature. At 20 °C the radiation is infrared. At 800 °C the radiation is red. At 3000 °C the radiation is white.
- The **Stefan-Boltzmann law** gives the rate of energy radiation from a surface

$$H = Ae\sigma T^4$$

where  $A$  is its surface area,  $e$  is a dimensionless constant between 0 and 1 called the **emissivity** of the surface (1 would be a perfect radiator),  $\sigma$  is the **Stefan-Boltzmann constant**

$$\sigma = 5.670\,374\,42 \times 10^{-8} \text{ W}/(\text{m}^2 \text{ K}^4),$$

and  $T$  is the temperature in Kelvin.

- An object's surroundings also emit radiation which is absorbed by the object. The net heat current from the object is this

$$H = Ae\sigma(T^4 - T_s^4)$$

where  $T_s$  is the temperature of the surroundings. At thermal equilibrium there is no heat flow.

- An object that is a good absorber must also be a good emitter. An ideal radiator with  $e = 1$  is also an ideal absorber, absorbing all the radiation that hits it. Such a surface is called an **ideal black body** or a **blackbody**.

## 18 Thermal Properties of Matter

### 18.1 Equations of State

- Quantities such as pressure, volume, temperature, and amount of substance are called **state variables**.
- An equation that relates state variables is called an **equation of state**.
- Molar mass (i.e. the mass of one mole of a substance) is sometimes called **molecular weight**.
- The **ideal gas equation** relates the absolute pressure  $p$ , volume  $V$ , number of moles  $n$ , and absolute temperature (in Kelvin)  $T$

$$pV = nRT$$

or

$$pV = \frac{m_{\text{total}}}{M}RT$$

or

$$pV = NkT$$

where  $R$  is the gas constant

$$R = 8.314 \text{ J}/(\text{mol K}).$$

- If the amount of a gas is constant, i.e.  $nR$  is constant, then

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2} = \text{constant}.$$

- **Standard temperature and pressure** (STP) is defined as  $0^\circ\text{C}$  and 1 atm.

- The **van der Waals** equation extends the ideal gas law to account for the effects of interactions between gas molecules and their finite size

$$\left(p + \frac{an^2}{V^2}\right)(V - nb) = nrT$$

where  $a$  and  $b$  are constants that vary between gases,  $a$  depends on the strength of attractive forces between the molecules, and  $b$  depends on the volume of the molecules.

- A  $pV$ -diagram is a two-dimensional diagram plotting pressure as a function of volume. Each line, called an **isotherm**, describes the relationship between pressure and volume at a particular temperature.
- In the  $pV$ -diagrams of non-ideal gases, isotherms can have flat areas. These represent periods where the gas is condensing into a liquid.
- The area under a  $pV$ -curve represents the work done by the system during a volume change.

## 18.2 Molecular Properties of Matter

- The force between molecules in a gas varies with  $r$ .
- An ideal gas is one whose molecules exert no attractive forces on each other and thus have no potential energy.
- The number of molecules in a mole is given by **Avagadro's number**

$$N_A = 6.022\,140\,76 \times 10^{23} \text{ molecules/mol.}$$

## 18.3 Kinetic-Molecular Model of an Ideal Gas

- The average translational kinetic energy of an ideal gas is

$$K_{\text{tr}} = \frac{3}{2}nRT.$$

- The **Boltzmann constant** is

$$k = \frac{R}{N_A} = 1.381 \times 10^{-23} \text{ J/(molecule K).}$$

- The average translational kinetic energy of a gas molecule is

$$\frac{1}{2}m(v^2)_{\text{av}} = \frac{3}{2}kT$$

and thus the **root-mean-square speed** of a gas molecule is

$$v_{\text{rms}} = \sqrt{(v^2)_{\text{av}}} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{M}}$$

where  $m$  is the mass of a molecule and  $M$  is the molar mass.

- The average distance travelled by a particle between collisions is called the **mean free path**

$$\lambda = \frac{V}{4\pi\sqrt{2}r^2N} = \frac{kT}{4\pi\sqrt{2}r^2p}$$

where  $V$  is the volume of the gas,  $r$  is the radius of a gas molecule,  $N$  is the number of molecules in the gas,  $T$  is its temperature, and  $p$  is its pressure.

## 18.4 Heat Capacities

- The **molar heat capacity at constant volume** of a particular gas is determined by its **degrees of freedom**, i.e. the number of ways in which it can store kinetic energy.
  - Monatomic gases have three degrees of freedom: one for each translational axis.
  - Diatomic gases have five degrees of freedom: one for each translational axis and two for rotational axes (the third is excluded because it's not affected by collisions).
  - Polyatomic gases have more than five degrees.
- The **principle of equipartition of energy** states that each degree of freedom contributes  $\frac{1}{2}kT$  to the average kinetic energy of the gas, i.e. for monatomic gases the average kinetic energy per molecule is  $\frac{3}{2}kT$ , for diatomic  $\frac{5}{2}kT$ , etc.
- The molar heat capacity at constant volume of a gas is equal to

$$C_V = \frac{d}{2}R$$

where  $d$  is the number of degrees of freedom.

- Vibrational energy can also contribute to the molar heat capacity at constant volume, but for most diatomic molecules this isn't the case.
- Molar heat capacity at constant volume is temperature dependent, with rotational and vibrational degrees of freedom only coming in to play at higher temperatures.
- Crystalline solids, i.e. solids whose atoms are arranged in a three-dimensional matrix, have six degrees of freedom: three translational and three from the potential energy of intermolecular forces. This results in the **rule of Dulong and Petit** which states that the molar heat capacity at constant volume of an ideal monatomic solid is

$$C_V = 3R.$$



## 18.5 Molecular Speeds

- In regards to the speeds of molecules in a gas, a **distribution function**  $f(v)$  gives the probability per unit speed of finding a particle with a speed near  $v$ . That is, the area under the curve between  $f(v_1)$  and  $f(v_2)$  gives the probability of finding a particle with speed in that range.
- The **Maxwell-Boltzmann distribution** is a particular distribution function

$$f(v) = 4\pi \left( \frac{m}{2\pi kT} \right)^{3/2} v^2 e^{-mv^2/2kT}$$

where  $m$  is the mass of a molecule,  $T$  is the absolute temperature of the gas, and  $v$  is the molecular speed.

- The peak of the curve and thus the most probable speed occurs at

$$v_{\text{mp}} = \sqrt{\frac{2kT}{m}}.$$

- The average speed is

$$v_{\text{av}} = \sqrt{\frac{8kT}{\pi m}}.$$

- The rms speed is

$$v_{\text{rms}} = \sqrt{\frac{3kT}{m}}.$$

## 18.6 Phases of Matter

- A  **$pT$  phase diagram** is a two-dimensional graph with pressure  $p$  on the vertical axis and temperature  $T$  on the horizontal axis. The plane is separated into three distinct regions: one for each phase of the material (solid, liquid, and gas).
- The borders between regions represent points of **phase equilibrium** where both phases can coexist at the same pressure and temperature.
- The **fusion curve** separates the solid and liquid regions.
- The **vaporisation curve** separates the liquid and gas regions.
- The **sublimation curve** separates the solid and gas regions.
- The three curves meet at the **triple point** — the only conditions under which all three phases can coexist.
- The vaporisation curve ends at the **critical point**. As the pressure and temperature of a substance approaches those of its critical point the physical differences between the liquid and gas phases decrease. At the critical

point there are no differences and a substance whose pressure or temperature is gradually decreased won't undergo a phase transition — instead its properties will continuously change from those of a gas to those of a liquid or vice versa.

## 19 The First Law of Thermodynamics

### 19.1 Thermodynamic Systems

- A **thermodynamic system** is any collection of objects that is convenient to regard as a unit, and that may have the potential to exchange energy with its surroundings.
- A process in which there are changes in the state of a thermodynamic system is called a **thermodynamic process**.
- In a thermodynamic process,  $Q$  represents the heat added to the system (increasing its energy) and  $W$  represents work done by the system on its surroundings (decreasing its energy). Both quantities may be negative in which case they represent heat being removed from the system (decreasing its energy) and the surroundings doing work on the system (increasing its energy), respectively.

### 19.2 Work Done During Volume Changes

- The work done during a volume change is

$$W = \int_{V_1}^{V_2} p dV$$

where  $V_1$  is the initial volume,  $V_2$  is the final volume, and  $p$  is the pressure.

- In the above equation  $p$  is often a function of  $V$ , in which case the integral represents the area under the graph of  $p = p(V)$ . This is the area under the curve of a  $pV$ -diagram.
- When volume increases, work is positive. When volume decreases, work is negative.

### 19.3 Paths Between Thermodynamic States

- When a thermodynamic system changes from an initial state to a final state it passes through a series of intermediate states. This series of states is called a **path**.
- There are an infinite number of paths between any given initial and final states, but the amount of work done by the system on its surroundings can differ between paths.

- The uncontrolled expansion of a gas into vacuum is called **free expansion**. The gas does no work as it expands so its temperature doesn't change.

## 19.4 Internal Energy and the First Law of Thermodynamics

- The **internal energy**  $U$  of a system is the sum of the kinetic energies of all its constituent particles, plus the sum of the potential energies of interactions between these particles.
- The **first law of thermodynamics** states

$$\Delta U = Q - W$$

where  $\Delta U$  is the change in internal energy of a thermodynamic system,  $Q$  is the heat added to the system, and  $W$  is the work done by the system on its surroundings.

- The change in internal energy of a thermodynamic system depends only on its initial and final states, not on the path taken.
- A **cyclic process** is one that starts and ends in the same state. In that case  $\Delta U = 0$  and thus  $Q = W$ : if a net amount of work was done by the system during this process, an equal amount of energy must have flowed into the system as heat  $Q$ .
- In an **isolated system** that does no work and experiences no heat flow,  $Q = W = 0$  and thus  $\Delta U = 0$ . This means that the internal energy of an isolated system is constant.
- In a cyclic process, the total work is positive if the process moves clockwise around the  $pV$ -diagram and negative if it moves counterclockwise.
- The first law of thermodynamics can also be expressed using infinitesimals

$$dU = dQ - dW = dQ - p dV.$$

## 19.5 Kinds of Thermodynamic Processes

- An **adiabatic process** is one with no heat transfer into or out of the system, i.e.  $Q = 0$ . This can be achieved by thermally insulating the system or performing the process so quickly that no heat transfer can occur. The path followed by an adiabatic process on a  $pV$ -diagram is called an **adiabat**.
- An **isochoric process** is one where the volume of the system doesn't change, i.e.  $W = 0$ . The path followed by an isochoric process on a  $pV$ -diagram is called an **isochor**.

- An **isobaric process** is one where the pressure of the system doesn't change, i.e.  $W = p(V_2 - V_1)$ . The path followed by an isobaric process on a  $pV$ -diagram is called an **isobar**.
- An **isothermal process** is one where the temperature of the system doesn't change. Any heat flowing into or out of the system must occur slowly enough to maintain thermal equilibrium. The path followed by an isothermal process on a  $pV$ -diagram is called an **isotherm**.

## 19.6 Internal Energy of an Ideal Gas

- The internal energy of an ideal gas depends only on its temperature, not on its pressure or volume.

## 19.7 Heat Capacities of an Ideal Gas

- Gases have two molar heat capacities: one at constant volume  $C_V$ , and one at constant pressure  $C_P$ .
- When heating a gas at constant pressure, the gas must do work on the container to increase its volume (otherwise pressure would increase). This means not all of the heat is used to increase the temperature of the gas. Thus, the amount of heat required to raise the temperature of a gas at constant pressure is greater than a gas at constant volume, i.e.  $C_P > C_V$ .
- For ideal gases

$$C_P = C_V + R.$$

- The **ratio of heat capacities** of a substance is denoted

$$\gamma = \frac{C_P}{C_V}.$$

- Because the internal energy of an ideal gas depends only on its temperature, not on its pressure or volume, if  $\Delta U = nC_V\Delta T$  is valid for one type of process (an isochoric or constant volume process), it must be valid for all other processes.

## 19.8 Adiabatic Processes for an Ideal Gas

- An adiabatic expansion results in a drop in temperature and an adiabatic compression results in a rise in temperature.

- For an adiabatic process

$$T_1 V_1^{\gamma-1} = T_2 V_2^{\gamma-1}$$

$$p_1 V_1^\gamma = p_2 V_2^\gamma$$

$$W = nC_V(T_1 - T_2)$$

$$\begin{aligned} W &= \frac{C_V}{R}(p_1 V_1 - p_2 V_2) \\ &= \frac{1}{\gamma - 1}(p_1 V_1 - p_2 V_2) \end{aligned}$$

## 20 The Second Law of Thermodynamics

### 20.1 Directions of Thermodynamic Processes

- Thermodynamic processes that occur between objects that aren't in thermodynamic equilibrium are irreversible, e.g. heat flowing from an area of higher to lower temperature, free expansion of gas, etc.
- Thermodynamic processes that occur between objects that are very nearly in thermodynamic equilibrium approach being reversible. They are also called **equilibrium processes**.

### 20.2 Heat Engines

- Any device that transforms heat into work or mechanical energy is called a **heat engine**.
- Usually a quantity of matter inside the engine undergoes inflow and outflow of heat, expansion and compression, and sometimes change of phase. We call this matter the **working substance** of the engine.
- All heat engines absorb heat from a **hot reservoir** and discard waste heat to a **cold reservoir**.
- The amount of heat absorbed from the hot reservoir is denoted  $Q_H$  (positive) and the amount of heat discarded to the cold reservoir is denoted  $Q_C$  (negative).
- If a heat engine undergoes a cyclic process  $\Delta U = 0$  so

$$W = Q = Q_H + Q_C = |Q_H| - |Q_C|.$$

- The **thermal efficiency** of a heat engine

$$e = \frac{W}{Q_H} = 1 + \frac{Q_C}{Q_H} = 1 - \left| \frac{Q_C}{Q_H} \right|$$

represents the fraction of input energy that's converted into useful work.

## 20.3 Internal Combustion Engines

- The **Otto cycle** is an idealised version of the thermodynamic process inside a gasoline engine.
- The thermal efficiency of the Otto cycle is

$$e = 1 - \frac{1}{r^{\gamma-1}}$$

where  $r$  is the compression ratio of the engine.

## 20.4 Refrigerators

- A fridge is like a reverse heat engine — it takes heat from a cold place (the inside of the fridge) to a hot place (the outside of the fridge). However, where a heat engine has a net output of work a fridge has a net input of work.
- When modelling a fridge as a thermodynamic system,  $Q_C$  is positive,  $Q_H$  is negative, and  $W$  is negative.
- A more performant fridge will maximise  $Q_C$  (the heat removed from the inside of the fridge) for a given  $W$  (the power input to the fridge). This is captured by the **coefficient of performance**

$$K = \frac{|Q_C|}{|W|} = \frac{|Q_C|}{|Q_H| - |Q_C|}.$$

- Another way of thinking about the performance of an air conditioner or fridge is in terms of the amount of heat removed per unit time, i.e. the heat current  $H$ , and the energy input per unit time, i.e. the power  $P$ . In terms of these units the coefficient of performance is

$$K = \frac{|Q_C|}{|W|} = \frac{Ht}{Pt} = \frac{H}{P}.$$

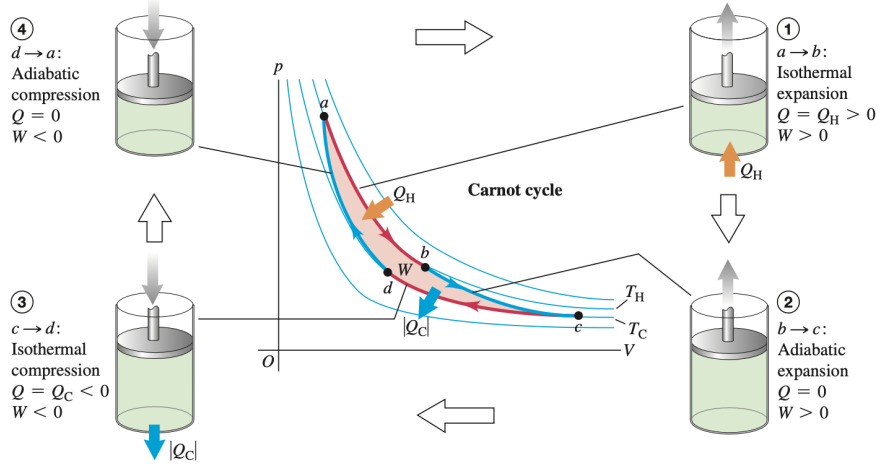
## 20.5 The Second Law of Thermodynamics

- The **second law of thermodynamics** states that it is impossible for any system to undergo a process in which it absorbs heat from a reservoir at a single temperature and converts the heat completely into mechanical work, with the system ending in the same state in which it began.
- An alternative wording is that it is impossible for any process to have as its sole result the transfer of heat from a cooler to a hotter object.

## 20.6 The Carnot Cycle

- The **Carnot cycle** is an ideal thermodynamic process that provides an upper limit on the maximum efficiency of an engine. In it, only reversible thermodynamic processes may be used — namely adiabatic and isothermal processes.

Figure 20.13 The Carnot cycle for an ideal gas. The light blue lines in the  $pV$ -diagram are isotherms (curves of constant temperature) and the dark blue lines are adiabats (curves of zero heat flow).



- The amount of heat transferred to/from the reservoirs in a Carnot cycle are related to the temperatures of the reservoirs by

$$\frac{Q_C}{Q_H} = -\frac{T_C}{T_H}.$$

- The efficiency of a Carnot engine is

$$e_{\text{Carnot}} = 1 - \frac{T_C}{T_H}.$$

- Because all processes in the Carnot cycle are reversible, the cycle itself can be reversed to make a Carnot fridge.
- The coefficient of performance of a Carnot fridge is

$$K_{\text{Carnot}} = \frac{T_C}{T_H - T_C}.$$

## 20.7 Entropy

- Entropy  $S$**  is a measure of the randomness in a system. When you add heat to a system the constituent particles gain kinetic energy, there is more randomness to their positions, and thus entropy increases. When you remove heat from a system entropy decreases.

- The units of entropy are J/K.
- The change in entropy during a reversible isothermal process is

$$\Delta S = \frac{Q}{T}.$$

- The change in entropy during any reversible thermodynamic process can be calculated as

$$\Delta S = \int_1^2 \frac{dQ}{T}$$

where  $\Delta S$  is the change in entropy, 1 is the initial state, 2 is the final state,  $dQ$  is the infinitesimal heat flow into the system, and  $T$  is the absolute temperature of the system.

- Entropy is determined entirely by the state of a system, not how it got there, i.e. entropy is path independent.
- In an adiabatic process  $Q = 0$  and thus there is no change in entropy.
- To calculate the change in entropy of an irreversible process, come up with a series of reversible processes that take the system between the same states and calculate the change in entropy of that.
- When a system undergoes a reversible change from state  $a$  to state  $b$ , the change in entropy is 0. This means that the change in entropy of adiabatic and isothermal processes are 0 (because they are reversible). The Carnot cycle is composed of adiabatic and isothermal processes so its change in entropy is also 0. In general, the change in entropy for any reversible cyclic process is 0.
- An alternative formulation of the second law of thermodynamics is that no process is possible in which total entropy decreases, when all systems that take part in the process are included.

## 20.8 Microscopic Interpretation of Entropy

- A description of the **macroscopic state** of a system describes the system as a whole rather than individual components of the system, e.g. “50% of the coins are heads, 50% are tails.”
- A description of the **microscopic state** of a system describes the individual components, e.g. “Coin 1 is heads, coin 2 is tails, etc.”
- Many microscopic states may correspond to a given macroscopic state.
- For any thermodynamic system, the most probable macroscopic state is the one with the greatest number of corresponding microscopic states, which is also the macroscopic state with the greatest randomness and the greatest entropy.



- The entropy of a given macroscopic state is

$$S = k \ln w$$

where  $k$  is the Boltzmann constant and  $w$  is the number of microscopic states associated with the macroscopic state.