3. Thermal Physics

Topic Outline

| Topic	Recommended Time	Giancoli Sections
3.1 Thermal Concepts	2h	14.1, 14.7-14.9, 15.1, 15.2, 15.4, 15.7
3.2 Thermal Properties	5h	14.4, 14.6, 14.3

3.1 Thermal Concepts

Temperature

- At a macroscopic (observable) level, **temperature** is the *average* random kinetic energy of the particles in an object

- At a microscopic level, temperature is the average random kinetic energy per molecule

- Temperature is a scalar quantity

Temperature Scales

- To set a temperature scale, you need two fixed points that can be reproduced under a variety of situations
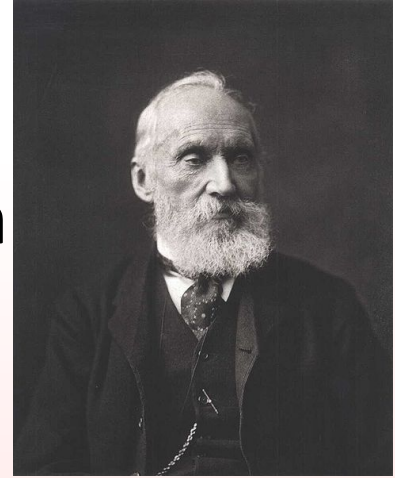
Temperature Scales
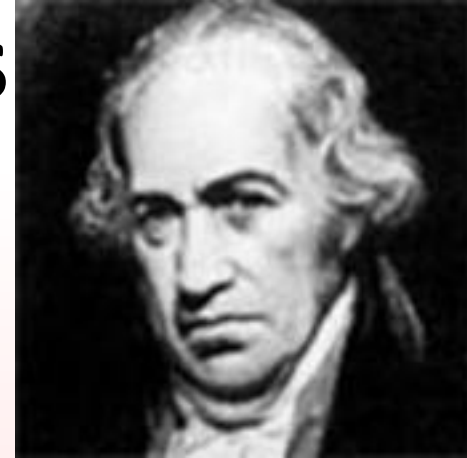
- The **Fahrenheit** temperature scale is named after the German physicist Daniel Gabriel Fahrenheit (1686-1736), who proposed the scale in 1724

- The zero fixed point was determined by putting a thermometer in a mixture of ice, water and ammonium chloride, which is a 'frigorific' mixture and stabilises at a particular temperature

- Other fixed points were the freezing point of water (32℉) and the boiling point of water (212℉), which were set 180° apart

Temperature Scales

- The **Celsius** scale is named after the Swedish astronomer Anders Celsius (1701-1744)

- It was originally based on the freezing point (0°C) and boiling point (100°C) of water at a pressure of one standard atmosphere (100 kPa)

Temperature Scales

- The **Kelvin** scale is named after the British physicist Lord Kelvin (1824-1907)
- Kelvin is the SI base unit for temperature
- **Absolute zero** (0 K = -273.15 ℃) is the (theoretical) point at which atoms have minimum kinetic energy
- The **triple point of water** is the temperature at which ice, liquid water and water vapour are all in equilibrium; this occurs at 273.16 K or 0.01 ℃
- Absolute zero and the triple point of water are now used as the fixed points for defining the Kelvin and Celsius temperature scales

Triple Point
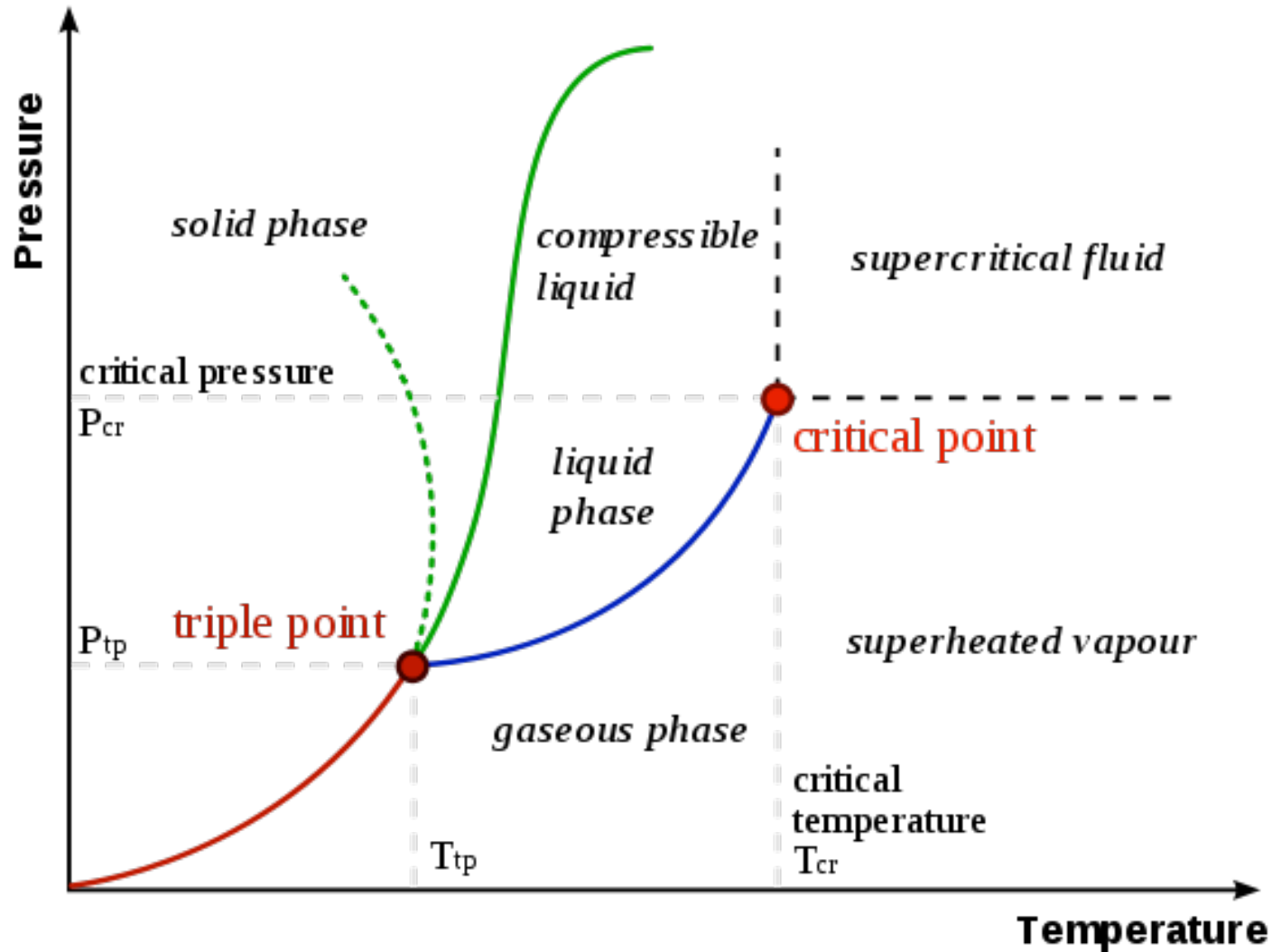
Thermometric Properties

- A **thermometric property** varies with temperature

- Examples of thermometric properties include:

  - the expansion of mercury with temperature

  - differing rates of expansion of two metals (bimetallic strips)

  - the pressure of a gas at constant volume

  - the volume of a gas at constant pressure

  - the electrical resistance of wires (resistors and thermistors)

  - the heating of two wires that are wound together (thermocouple)

  - the colour of a solid heated to high temperatures (pyrometers)
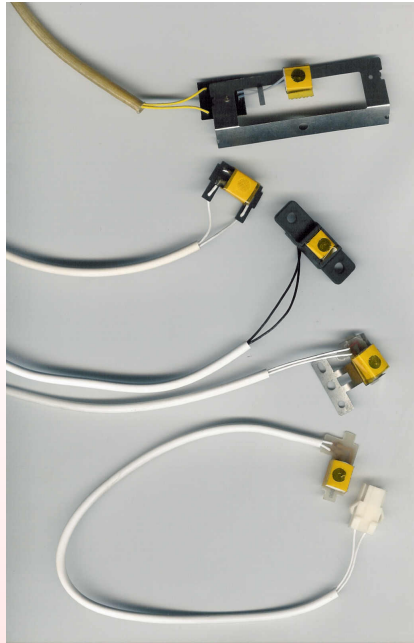
Thermal Energy

- **Thermal energy** is the *total* random kinetic energy of the particles in an object

- Thermal energy is measured in Joules

- The **random kinetic energy** of the particles includes:

  - vibrational energy
  - rotational energy
  - translational energy

Transfer of Thermal Energy

- **Heat** is the non-mechanical transfer of energy, i.e. the thermal energy that is absorbed, released or transferred from one object to another

- Heat is measured in Joules

- Thermal energy can be transferred as:
  - **Heat**: thermal energy of one object → thermal energy of another object
    - e.g. from boiling water to a ceramic mug
  - **Work**: thermal energy → mechanical energy
    - e.g. thermal energy of water vapour → the water vapour expands → the expansion drives a turbine.

Internal Energy

- The **internal energy** of an object is the sum of:
  - the random kinetic energy (thermal energy) of the particles
  - the potential energy *within* particles (nuclear and chemical potential energy)
  - the potential energy *between* the particles
- The potential energy between the particles is a result of the attractive (electromagnetic) force between particles
- Energy is needed to overcome this force and move the particles further apart (e.g. to go from solid → liquid)
- So when particles are further apart, they have more potential energy

Thermodynamics

- **Thermodynamics** is the study of:
  - the transfer of thermal energy (transfer via heat)
  - the conversion of thermal energy to other forms of energy (work done)
  - the relationship between thermal energy, temperature, pressure and volume

Zeroth Law of Thermodynamics

If two thermodynamic systems are each in thermal equilibrium with a third, then they are in thermal equilibrium with each other
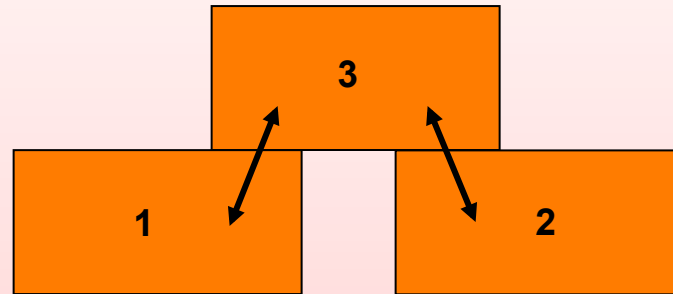
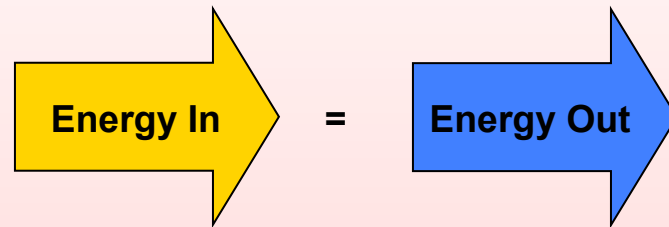- When two objects are in thermal equilibrium, there is no net transfer of thermal energy

- The zeroth law of thermodynamics is an example of an equivalence relationship

First Law of Thermodynamics

Energy is neither created or destroyed, but only changed from one form to another

Energy In = Energy Out

- If thermal energy is used to do work, the work done will be equal to the thermal energy supplied to the system

James Prescott Joule
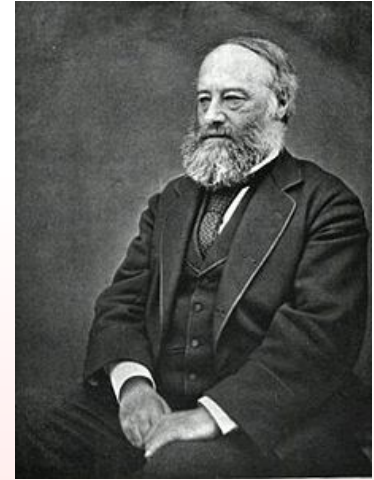
- James Prescott Joule (1818-1889) did a famous experiment to determine the mechanical equivalent of heat

- As the mass falls (losing gravitational potential energy) the paddles spin in the water, increasing the temperature of the water
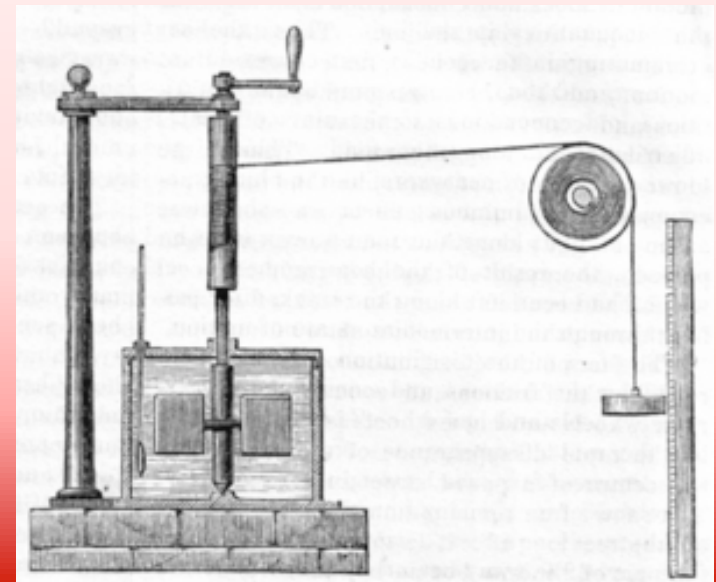
Second Law of Thermodynamics

The entropy of an isolated system that is not in equilibrium will tend to increase over time

(The entropy of the system will reach a maximum when the system is in equilibrium)

OR **Entropy tends to increase**

OR **Heat tends to go from a hot to a cold object**

- **Entropy** is a measure of the amount of energy in a system that is *not* available to do useful work

Entropy

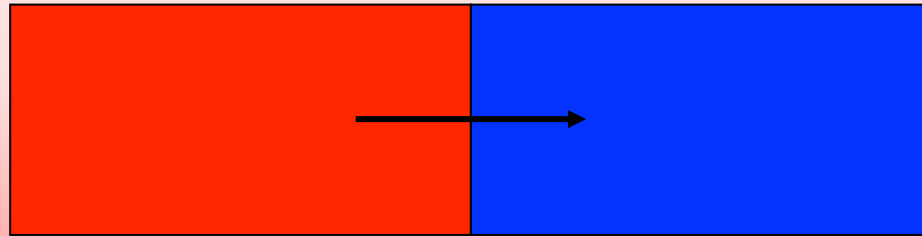Thermal energy can be used to do work, e.g. thermal energy from the combustion of petrol and oxygen can move a piston. But in systems in which energy is transferred, some energy is usually 'lost' as heat, e.g. a motor running loses some energy as heat through friction; a cup of hot coffee loses energy as heat to the surrounding environment.

When energy is lost as heat, that energy goes to random movements of particles in the environment. So energy is conserved, but it is very hard to get any 'useful' mechanical work out of the random movements of these particles, i.e. entropy has increased.

Entropy
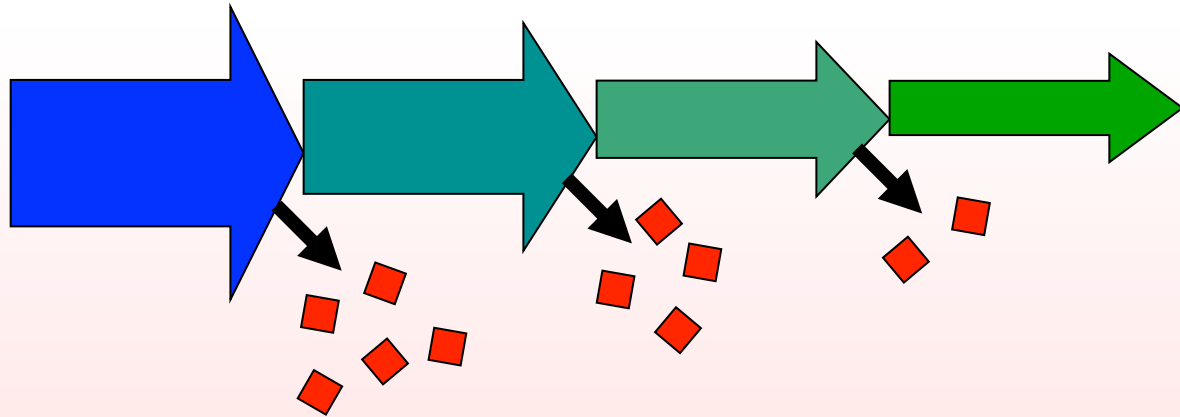
A series of energy transforms. At each stage, some energy is converted to thermal energy (random movement of particles in the environment). It is very hard to get useful work out of the random movement of these particles, and we say entropy has increased.

Third Law of Thermodynamics

As temperature approaches absolute zero, the entropy of a system approaches a minimum
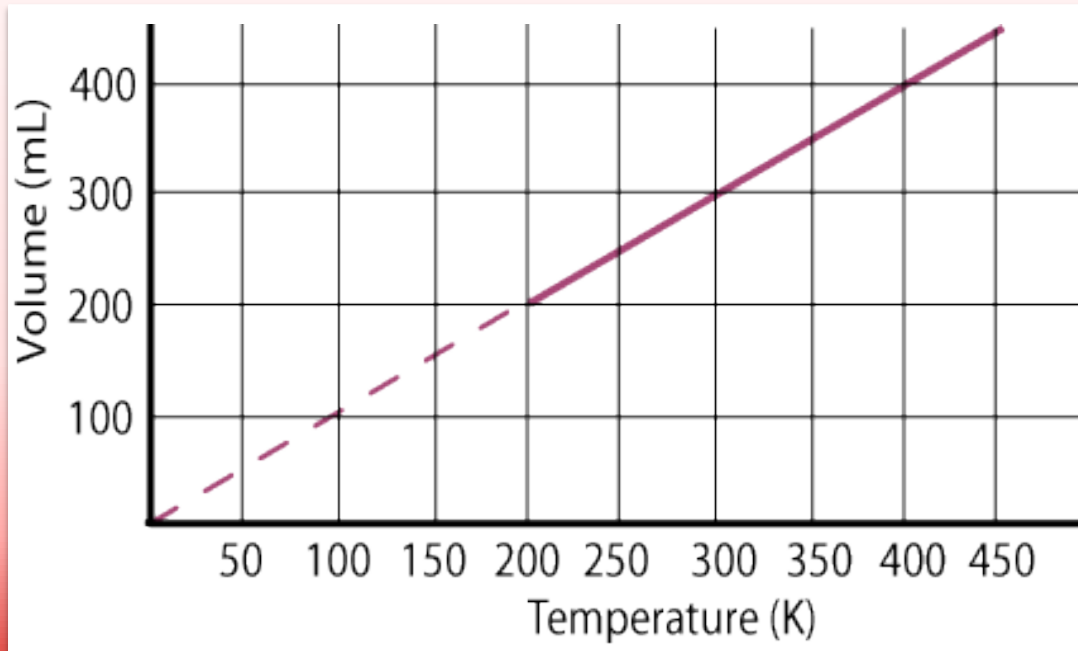
Transfer of Thermal Energy

- Thermal energy can be transferred by convection, conduction or radiation

Conduction

- **Thermal conduction** is the transfer of thermal energy from a hotter region to a colder region by particle collision without net movement of the substance
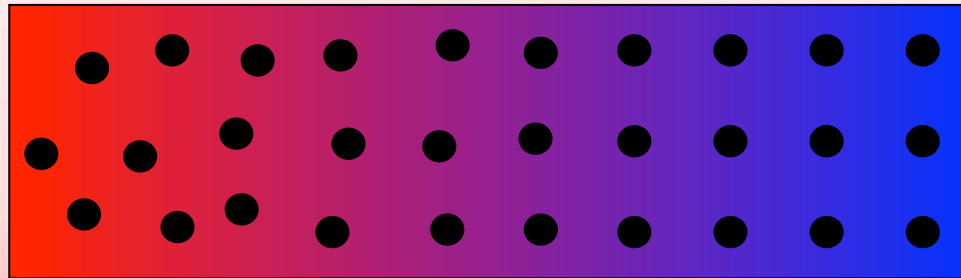
- Metals are good thermal conductors because they are dense (the particles are close together) and because the free electrons transfer vibrational energy readily

Convection

- **Thermal convection** is the bulk movement of particles from a region of high thermal energy to a region of low thermal energy
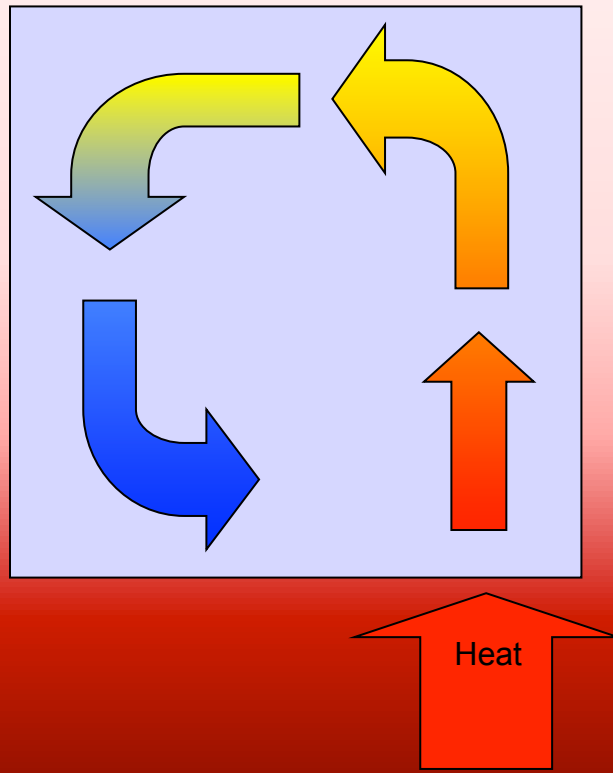
Heat

Convection

- Convection only occurs in liquids and gasses
- As the fluid is heated, it expands and becomes less dense, and so it rises in relation to the surrounding fluid; cooler, denser fluid, then replaces the heated fluid, resulting in the formation of a convection current
- A pump or a fan can also create convection (forced convection)

Radiation

- **Thermal radiation** is the transfer of heat by electromagnetic waves

- Radiation occurs independently of particles
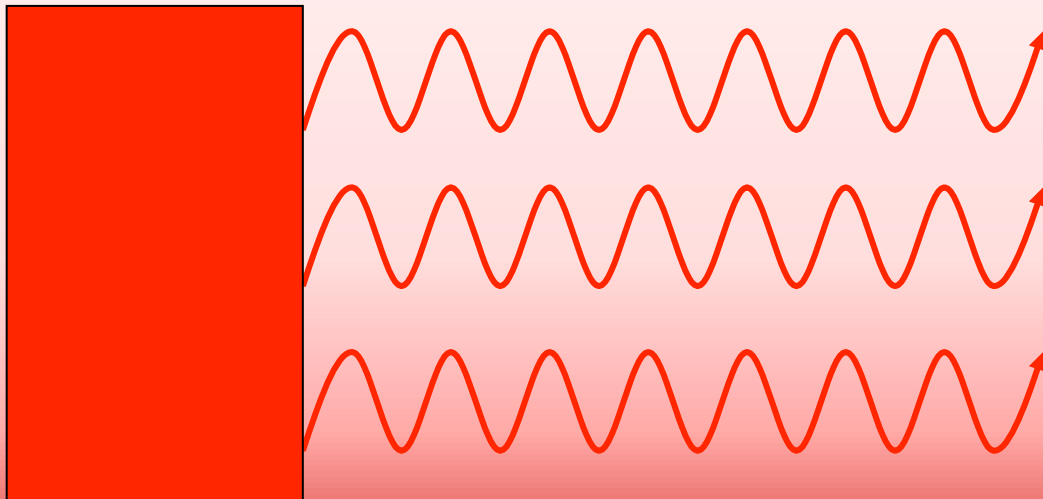
The Mole

- In 1961, the International Union of Pure and Applied Chemistry (IUPAC) defined the mass of atoms relative to the mass of carbon-12. C-12 was assigned the mass of 12.0000 atomic mass units, so a single proton or neutron has a mass of approximately 1 a.m.u.

- One **mole** of a substance contains 6.02×10^{23} particles, the same number of particles as there are in 0.012 kg of carbon-12. This number is also called **Avogadro's number** (N_A), and is named after the Italian physicist Amadeo Avogadro (1776-1856)

- The mole is the SI base unit for amount of substance

- The **molar mass** (M) of a substance is in g.mol^{-1}

$$m = nM$$

m = mass, g

n = moles, mol

M = molar mass, g.mol^{-1}

3.1 Thermal Properties of Matter

Heat Transfer

- According to the second law of thermodynamics, heat is transferred from hotter objects (higher temperature) to colder objects (lower temperature)

- There are three possible effects of transferring heat to a substance:

  1. Temperature rise
  2. Change of state
  3. Expansion

1. Temperature Rise

- **Temperature rise** occurs when heat is transferred to an object and the random kinetic energy of the particles increases

Heat Capacity

- Different substances heat up and cool down more quickly – this is related to how much thermal energy a substance can store

- Substances with a **high heat capacity** (high ability to store heat) will change temperature **slowly**; a large change in thermal energy is needed to bring about a change in temperature

- Substances with a **low heat capacity** will change temperature **more quickly**; a small change in thermal energy results in a large change in temperature

- For example, slices of tomato or pineapple on a pizza stay hot a lot longer than the pizza base. Tomato and pineapple have a higher heat capacity than the pizza base

Heat Capacity

- The **heat capacity** (or **thermal capacity**) of an object is the amount of energy needed to raise the temperature of that object by 1℃

- Heat capacity is measured in J.K^{-1}

$$Heat\ capacity = \frac{\Delta Q}{\Delta T}$$

ΔQ = change in thermal energy, J

ΔT = change in temperature, K

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 a substance by 1℃

- Specific heat capacity is measured in J.kg^{-1}.K^{-1}

$$c = \frac{\Delta Q}{m\Delta T}$$

c = specific heat capacity, J.kg^{-1}.K^{-1}

ΔQ = change in thermal energy, J

m = mass, kg

ΔT = change in temperature, K

Some Specific Heat Capacities

Substance	Phase	Specific Heat Capacity ($J \cdot kg^{-1} \cdot K^{-1}$)
Hydrogen	gas	14.3×10^3
Helium	gas	5.19×10^3
Water (liquid, 25℃)	liquid	4.18×10^3
Ethanol	liquid	2.44×10^3
Water (ice)	solid	2.1×10^3
Air	gas	1.01×10^3
Aluminium	solid	0.897×10^3
Diamond	solid	0.501×10^3
Iron	solid	0.450×10^3
Copper	solid	0.385×10^3
Mercury	solid	0.140×10^3
Gold	solid	0.129×10^3

- Worksheet: Heat Transfer Questions

- Practical: Heat 200 ml of oil and 200 ml of water to the same temperature, measure the temperature of each at 1 min intervals as the cool down, plot a graph of the results

- Practical: Determine the specific heat capacity of water using a calorimeter

2. Change of State

- **Kinetic theory** is used to explain the states of matter
- The assumptions of this theory are:
  - all matter is composed of extremely small particles
  - all particles are in constant motion
  - kinetic energy is conserved in collisions between these particles
  - a mutually attractive force exists between the particles
- **Brownian motion** refers to the random movement of particles in a fluid. It is observed when, e.g. pollen grains are scattered on a dish of water. This motion increases when the fluid is heated. Brownian motion is considered evidence of kinetic theory.

States of Matter

- There are four states of matter: solid, liquid, gas and plasma

- Plasma forms when the temperature of a gas is increased (or the pressure is decreased) and the atoms and molecules ionise

- Most of the matter in space is in the plasma state

States of Matter

	Solid	Liquid	Gas
Shape	definite	variable	variable
Volume	definite	definite	variable
Compressibility	very low	low	high
Diffusion	very slow	slow	fast
Density	high	high	low
Particles per m^3	10^{28}	10^{28}	10^{25}
Kinetic energy	vibrational only	Vibrational, rotational, some translational	Higher vibrational, higher rotational, v. high translational
Potential energy	high	higher	highest
Mean molecular separation	r_0	r_0	$10r_0$
Thermal energy	$<\varepsilon/10$	$\varepsilon/10<\ <\varepsilon$	$\varepsilon<$

r_0 = separation distance at which particles are at equilibrium, m

ε = difference potential energy between a solid and particles with infinite separation, J

Change of State

- A **change of state** occurs when heat is transferred to or from a substance and the potential energy of the particles changes, but the average kinetic energy (temperature) does not

- A change of state can occur when:
  - Heat is transferred to a substance and this energy is used to overcome attractive forces between the particles; as a result, the particles gain potential energy and move further apart, e.g. melting, evaporation, sublimation
  - Heat is transferred from a substance and the particles lose potential energy and move closer together, e.g. condensing, solidifying (freezing)

- Changes of state occur at fixed combinations of temperature and pressure for different substances, e.g. water melts at 0°C and boils at 100°C at standard pressure

- Iodine and carbon dioxide are examples of substances that sublime at standard pressure

Latent Heat

- **Latent heat** is the intermolecular potential energy that is gained or lost when a substance undergoes a change of state (phase change)

- Latent heat is measured in $J.kg^{-1}$

- The **latent heat of fusion** is the amount of energy required to melt 1 kg of a substance (or the amount of energy released when 1 kg of that substance solidifies)
  - The latent heat of fusion for water is 3.33×10^5 $J.kg^{-1}$

- The **latent heat of vaporisation** is the amount of energy required to evaporate 1 kg of a liquid (or the amount of energy released when 1 kg of the gas condenses)
  - The latent heat of vaporisation for water is 2.26×10^6 $J.kg^{-1}$

Latent Heat

$$\Delta Q = mL$$
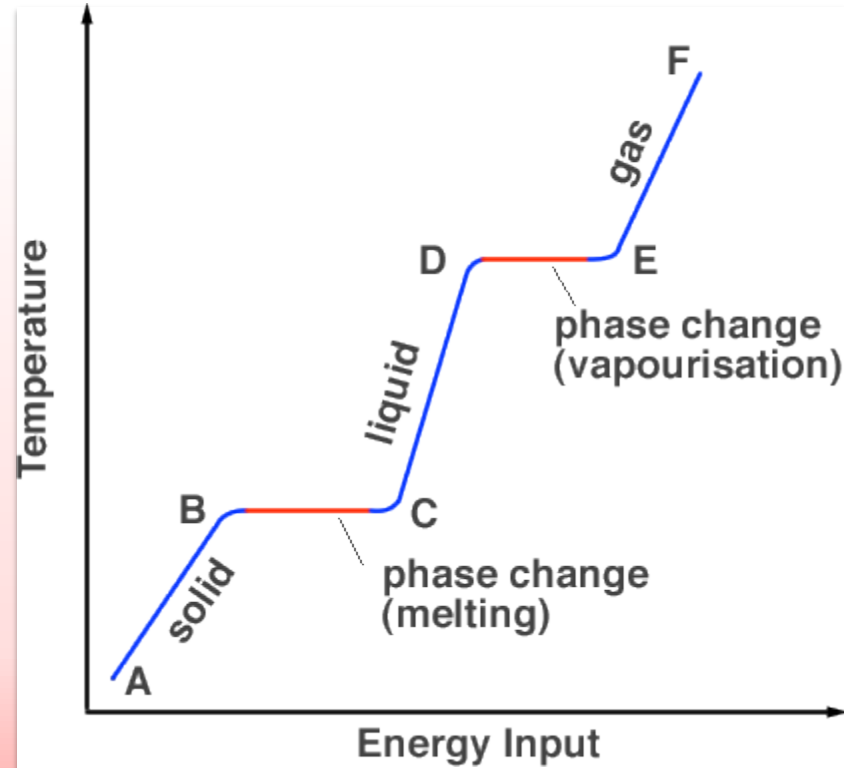
ΔQ = thermal energy gained or lost by the substance, J

m = mass, kg

L = latent heat, J.kg^{-1}

- Practical: get 3 thermometers; place dry cotton wool on one, cotton wool soaked in water on the second, and cotton wool soaked in meths on the third; record the temperature every 2 min for 10 min

Vapour Pressure

- Because temperature is related to the *average* kinetic energy of particles in a substance, some particles will escape a liquid (i.e. enter the gas phase) when the substance is below the boiling point

- The **boiling point** of a liquid is the temperature at which particles will enter the gas phase at a given pressure

- The **equilibrium vapour pressure** is the pressure at which particles will enter the gas phase at a given temperature

- The vapour pressure of a liquid gives an indication of how readily particles of a substance can enter the gas phase

- A **volatile liquid** evaporates readily and has a high vapour pressure

Vapour Pressure

- Some vapour pressures at 293K:

Substance	Vapour Pressure (kPa)
Ether	58.9
Chloroform	19.3
Ethanol	5.8
Water	2.3
Mercury	0.0002

For example, when the lid is on a carbonated drink (e.g. lemonade), the pressure in the bottle is greater than the vapour pressure for carbon dioxide and the carbon dioxide stays in solution. When you take the lid off the bottle, the pressure in the bottle drops to atmospheric pressure and the carbon dioxide starts to enter the gas phase.

3. Expansion

- <u>Expansion</u> occurs when a substance is heated
- The particles gain more kinetic energy and potential energy, i.e. they move further apart
- In general, gasses expand the most and solids expand the least

Kinematic Model of an Ideal Gas

- We use the molecular model of an **ideal gas** to study the relationship between the macroscopic properties (temperature, pressure, volume) and microscopic properties (interactions of particles) of a gas

- In an ideal gas, the particles move in random straight lines until they collide with each other or with the walls of the container, resulting in an overall zig-zag pattern of motion

- Because the particles move freely and are relatively far apart, they occupy the total volume of the container

- The behaviour of most gasses approximates an ideal gas at low pressures

Kinetic Model of an Idea Gas

- The molecular model of an ideal gas has a number of assumptions:
  1. The gas is a pure substance (it consists of only one type of particle)
  2. The number of particles in the sample is large
  3. The particles are in constant random motion
  4. The range of the attractive forces between the particles is small compared to the average separation of the particles
  5. The size of the particles is small compared to the average separation of the particles
  6. The particles make perfectly elastic collisions with each other and the walls of the container
  7. These collisions are short in duration
  8. The particles obey Newton's laws of motion

Pressure

- **Pressure** is the force exerted over an area
- The unit of pressure is Pascals, Pa

$$P = \frac{F}{A}$$

P = pressure, Pa

F = force, N

A = area, m^2

Pressure

- In terms of the molecular model of an ideal gas, pressure is understood as the force exerted by the particles as they collide against the walls of the container

- The pressure of a gas is proportional to the number of particles per unit volume and to the average translational kinetic energy of the particles, i.e. there is a relationship between pressure, volume and temperature

IDEAL-GAS LAW

gas pressures

gas molecules

volume 1

volume 2
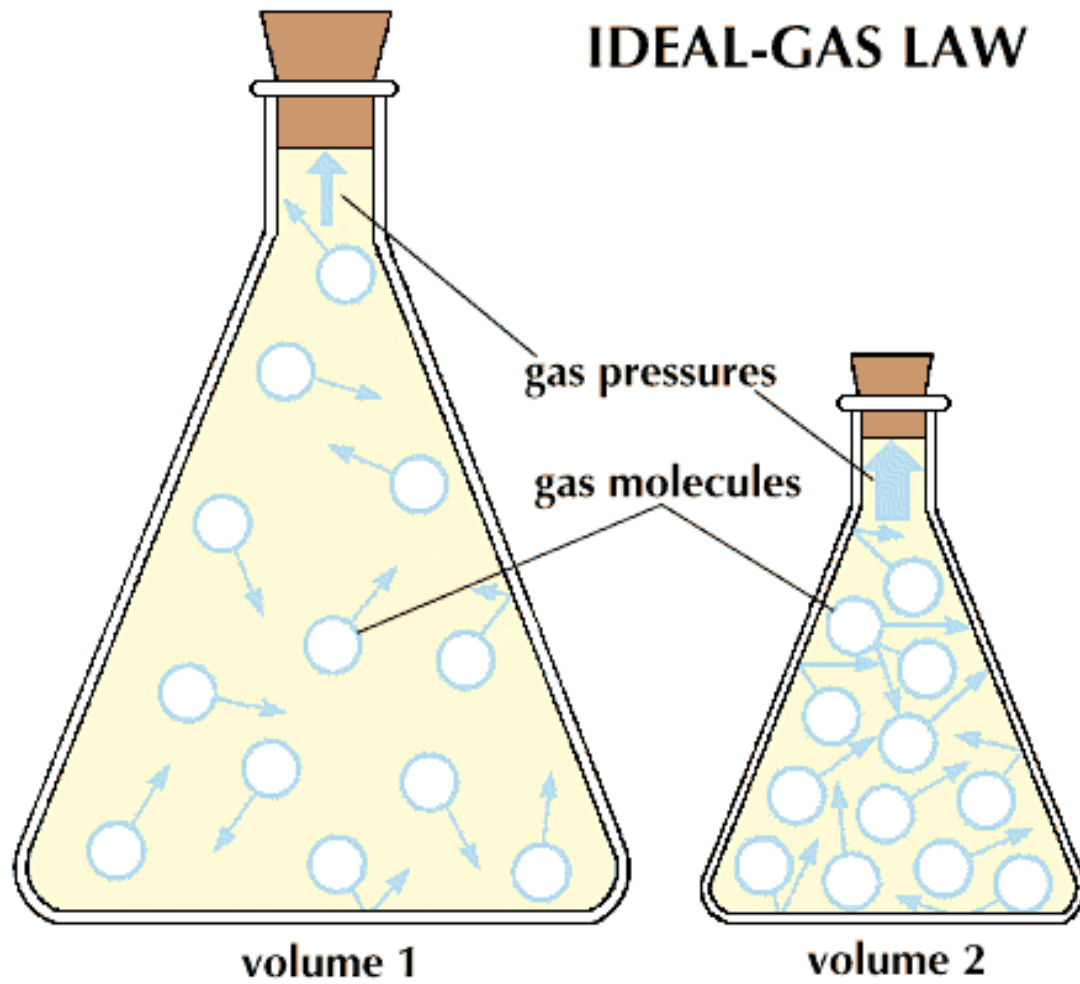
Atmospheric Pressure

- In 1643, Torricelli found that one atmosphere of pressure could support a vertical column of mercury 760 mm high

- In 1646, Pascal found that one atmosphere of pressure could support a vertical column of water 10.4 m high

- Standard atmospheric pressure is 101.3 kPa (or 1.01×10^5 Nm^{-2} or 760 mmHg)
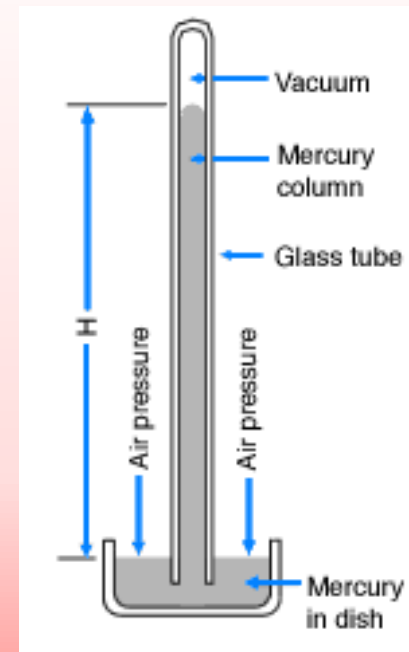
Temperature in an Ideal Gas

- At a macroscopic level, temperature is the average random kinetic energy of the particles

- At a microscopic level, temperature is the average kinetic energy per particle

- There is a broad range of speeds represented in the particles in a gas. The speed of the particles can change when they collide with each other

- The **Maxwell-Boltzmann speed distribution** shows the number of particles at different speeds for a gas at a given temperature. The peak of the graph shows the most probable speed of a particle in a gas of a given temperature.

Maxwell-Boltzmann Speed Distribution
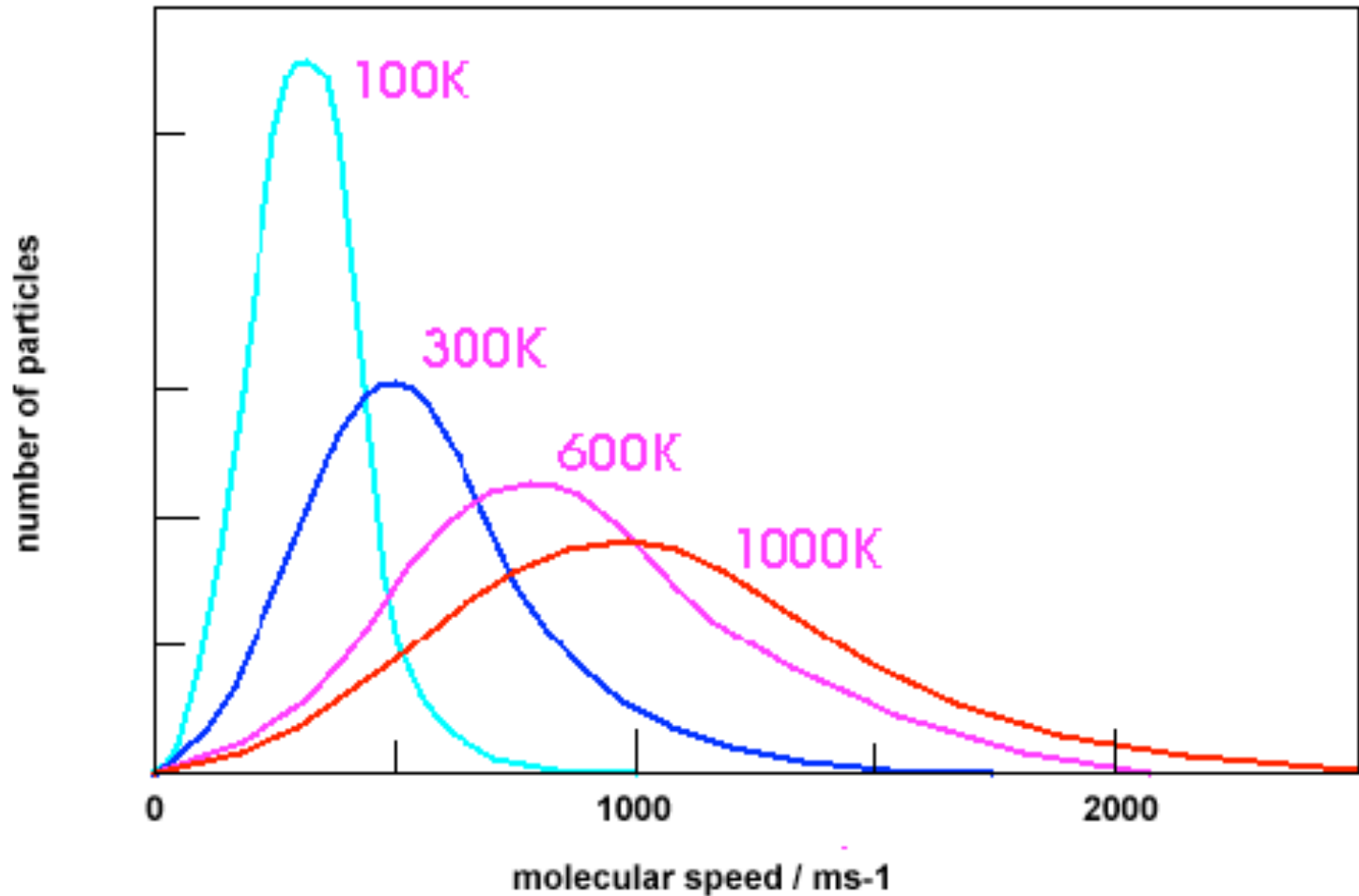
Kinetic Energy in an Ideal Gas

- A similar graph can be drawn for number of particles vs. kinetic energy of the particles at a given temperature

- For this graph, the area under the curve (i.e. number of particles x kinetic energy of the particles) gives the total kinetic energy of the gas

- In an ideal gas, the average kinetic energy per molecule is:

$$E_{Kav} = \frac{1}{2} mv_{av}^2 = \frac{3}{2} kT$$

- So the average speed squared is proportional to the temperature of the gas

E_K = average kinetic energy

m = mass of the particle, kg

v = average speed of the particles, $m.s^{-1}$

k = Boltzmann constant, 1.38×10^{-23} $J.K^{-1}$
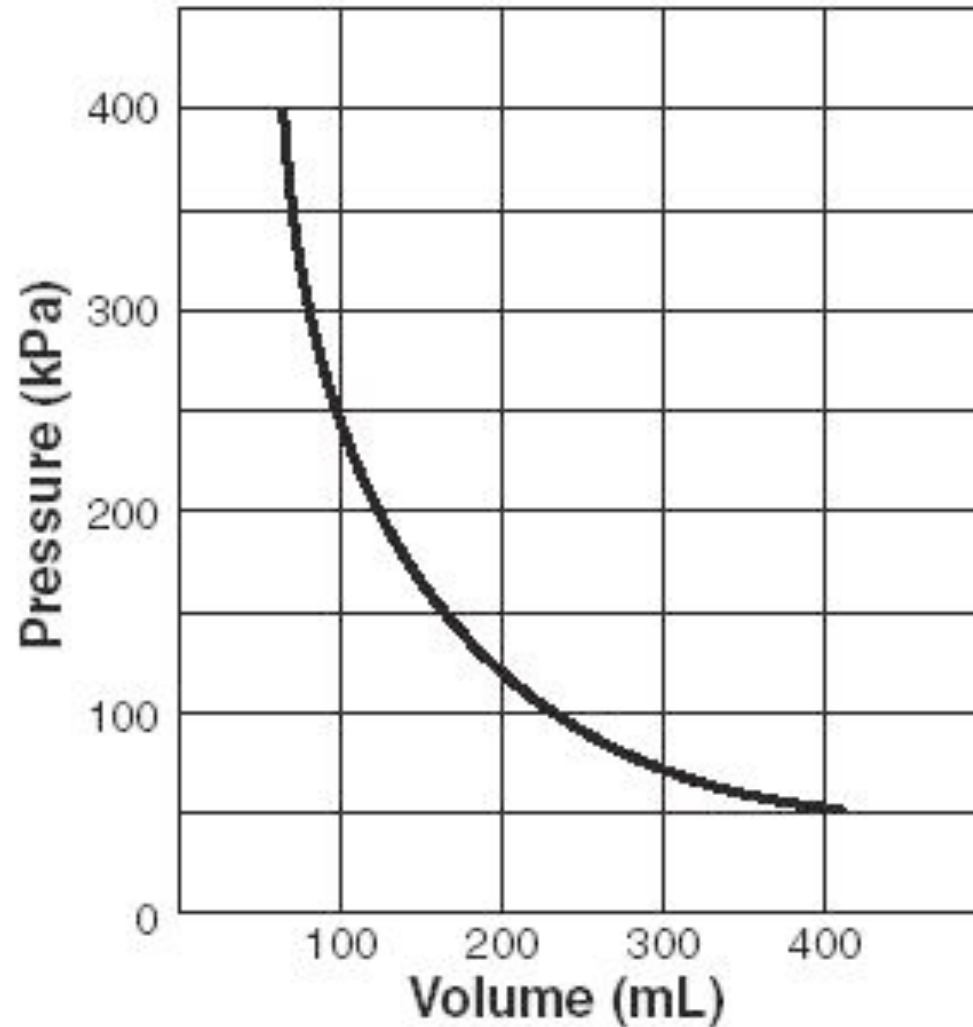
T = temperature of the gas, K

Boyle's Law

- Robert Boyle (1627-1691) investigated the relationship between pressure and volume of a gas at a fixed temperature

- He found that pressure and volume were inversely related, i.e. if pressure is increased, volume is decreased

$$P \propto \frac{1}{V}$$

- Plotting pressure vs volume (with temperature and amount of gas kept constant) will give an inverse relationship

Boyle's Law

Boyle's Law

- Another way of writing this relationship is that pressure multiplied by volume gives a constant

$$PV = \text{a constant}$$

- So if the amount of gas (in mol) and the temperature of the gas are kept constant, changes in pressure and volume can be worked out using the following relationship:

$$P_1V_1 = P_2V_2$$

P_1 = pressure during condition 1, Pa
V_1 = volume during condition 1, m^3
P_2 = pressure during condition 2, Pa
V_2 = volume during condition 2, m^3

Charles' (Gay-Lussac) Law

- Jacques Charles (1746-1823) investigated the relationship between temperature and volume of a gas

- Gay-Lussac (1778-1850) later published more accurate investigations

- Charles found that the volume of a gas (at constant pressure) is directly proportional to its absolute (Kelvin) temperature
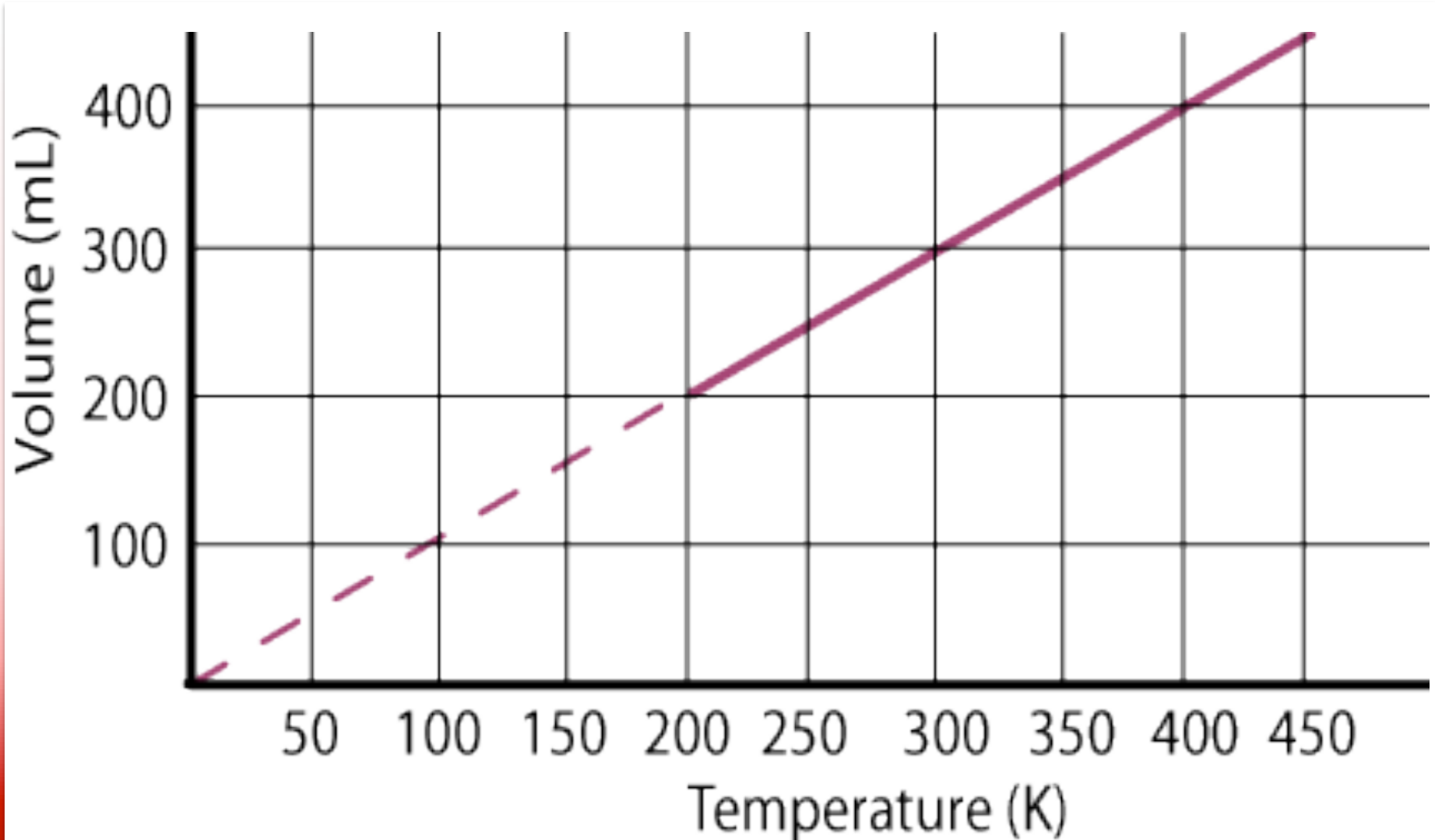
$$V \propto T$$

- This can also be written:

$$\frac{V}{T} = a\ constant$$

- One consequence of this law is that at a temperature of 0 K, the gas has a theoretical volume of 0 m^3

Charles' Law

Charles' Law

- So if the amount of gas (in mol) and the pressure of the gas are kept constant, changes in temperature and volume can be worked out using the following relationship:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V_1 = volume during condition 1, m^3

T_1 = temperature during condition 1, K

V_2 = volume during condition 2, m^3

T_2 = temperature during condition 2, K

The Pressure Law

- The states that the pressure of a gas (at constant volume) is directly proportional to its absolute (Kelvin) temperature
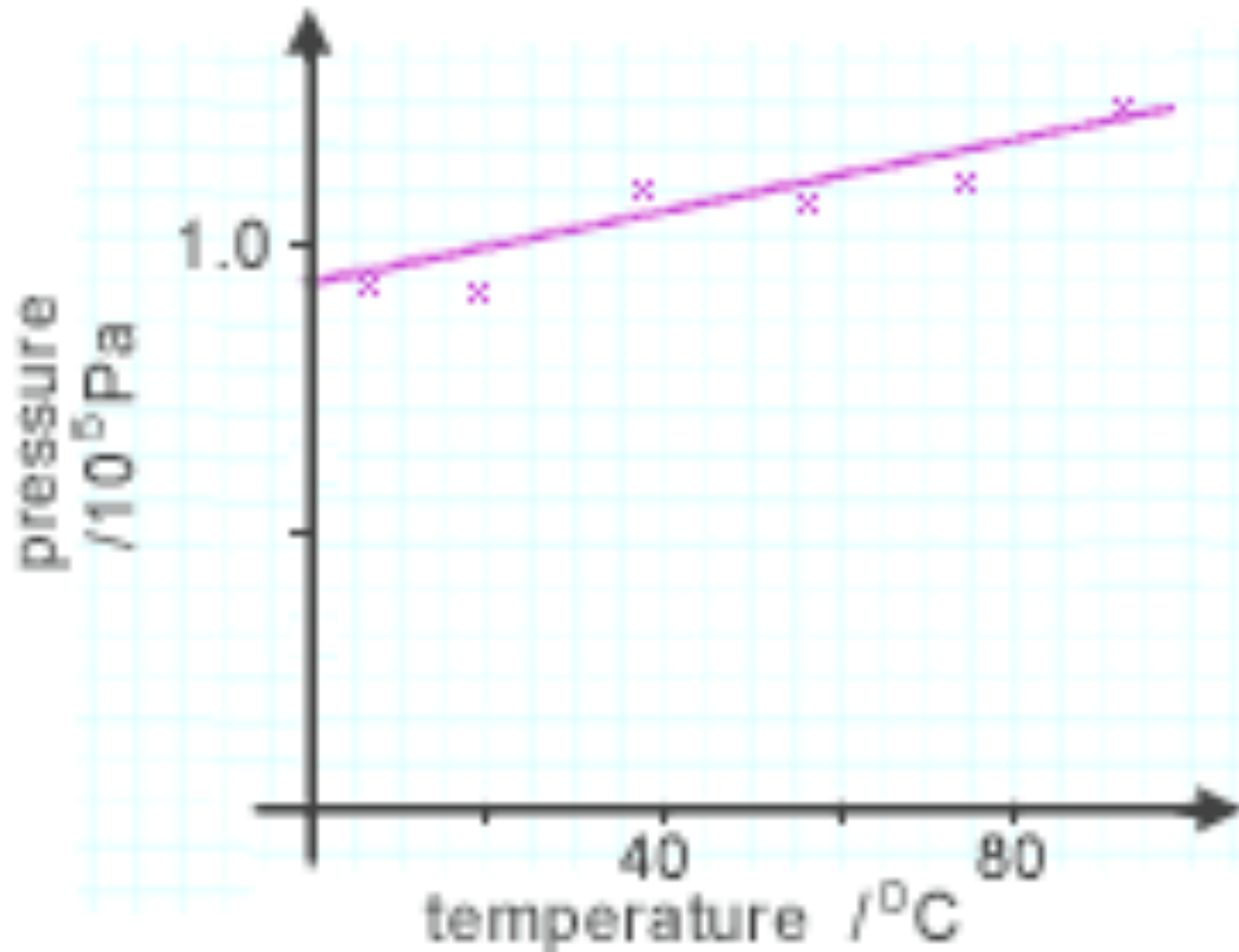
$$P \propto T$$

- This can also be written:

$$\frac{P}{T} = a\ constant$$

- One consequence of this is that at a temperature of 0 K, the gas has a theoretical pressure of 0 Pa

The Pressure Law

Pressure Law

- So if the amount of gas (in mol) and the volume of the gas are kept constant, changes in temperature and pressure can be worked out using the following relationship:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

P_1 = pressure during condition 1, Pa

T_1 = temperature during condition 1, K

P_2 = pressure during condition 2, Pa

T_2 = temperature during condition 2, K

The Ideal Gas Law

- The Ideal Gas Law (or Charles-Boyle's Law) combines Boyle's Law, the law of Charles and Gay-Lussac, and the Pressure Law

- It considers the relationship between pressure, volume, temperature and amount of gas
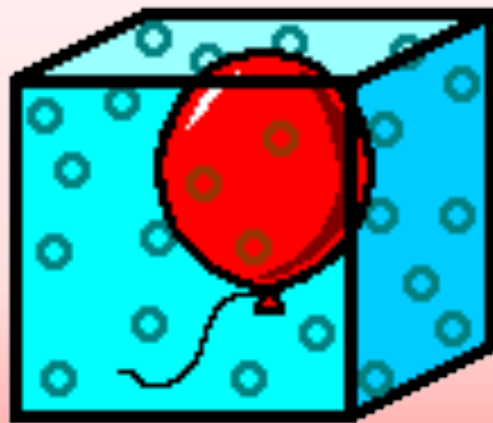
$$PV = nRT$$

P = pressure, Pa

V = volume, m^3

n = amount of gas, mol

R = universal gas constant, 8.3 $J.mol^{-1}.K^{-1}$

T = temperature, K
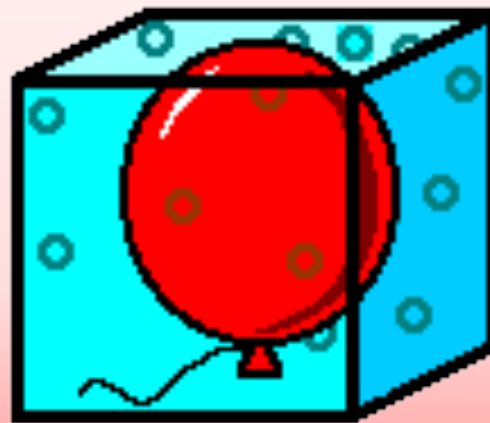
Figure 2. Volume of One Mole of Gas Under Different Conditions

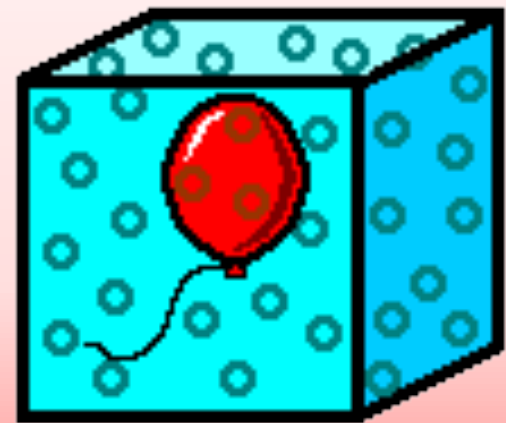All Balloons contain one mole of gas (6.02×10^{23} molecules)

A
T = Medium
P = Medium

B
T = High
P = Low

C
T = Low
P = High