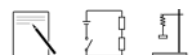# OCR B Physics – H557

Module 5: Rise and Fall of the Clockwork Universe

You should be able to demonstrate and show your understanding of:	Progress and understanding:			
	1	2	3	4
5.2: Matter				
Particle model: A gas consists of many very small, rapidly moving molecules that exert pressure by rebounding off the walls of the container. The force exerted by the gas depends on the number of particles colliding, their mass and their velocity				
Ideal gas: A gas where the molecules do not interact at all and occupies a negligible volume. - All gases behave as ideal gases when the molecules are well separated, high temperatures and low pressures. - Only small molecules such as hydrogen and helium still behave as ideal gases if the separation is small				
Assumptions for an ideal gas: -All molecules are identical (behave the same way) -The internal energy is entirely kinetic -The collisions are elastic (kinetic energy and momentum conserved) -The gas takes up a negligible volume -Newton's laws apply to the motion of molecules -The motion of molecules is random -Molecules all travel in straight lines between collisions (gravitational and electrostatic forces have no affect of the gas)				
Boyle's Law: Pressure of an ideal gas is inversely proportional to the volume of it, if temperature and amount of gas are held constant $$pV = k \qquad p_1V_1 = p_2V_2$$ - A P-V graph demonstrating Boyle's Law looks similar to a 1/x graph in the positive quadrant. If temperature is increased the graph moves away from the coordinate axis (think of how 1/x changes when the numerator is made larger) - If V is halved, the molecules strike the walls twice as often, doubling the pressure which follows Boyle's Law (the same effect is achieved by doubling the number of molecules, N, in the same original volume)				
Amount Law: Pressure of an ideal gas is directly proportional to the amount of gas, if temperature and volume are held constant. A p-N graph would be linear				

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Charles' Law: Volume of an ideal gas is directly proportional to temperature in Kelvin, if pressure and mass of gas are held constant $$\frac{V}{T} = k \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$ -A V-T graph demonstrating Charles' Law is linear. For temperature in kelvin the graph passes through the origin, in Celsius it has an x-intercept at x ≈ -273				
Pressure Law: Pressure of an ideal gas is directly proportional to temperature in Kelvin, if volume and mass of gas are held constant $$\frac{p}{T} = k \qquad \frac{p_1}{T_1} = \frac{p_2}{T_2}$$ -A P-T graph demonstrating the pressure law is linear. For temperature in kelvin the graph passes through the origin, in Celsius it has an x-intercept at x ≈ -273 -The pressure law predicts zero pressure for zero Kelvin				
These ideal gas laws do not apply when real gases are compressed by large amounts as the gas molecules interact with each other, exerting interactive Van der Waals/intermolecular forces. This also applies for high P, low V, or at T near boiling point because the atoms occupy a significant proportion of the volume				
Kelvin scale: 0K = -273.15 Celsius. Temperature change in kelvin is the same as the change in Celsius. Celsius to kelvin, +273.15. Kelvin to Celsius, -273.15. The Kelvin scale is related to the average kinetic energy per molecule (as it is measured form absolute zero) unlike the Celsius scale				
At very low temperatures (close to absolute zero) gases cool and contract so the molecules become close enough to interact before condensing to a liquid in some cases. So, at low temperatures, the reliability of the ideal gas laws falters				
Ideal gas equation: 1) Combining Boyle's ($p \propto \frac{1}{V}$) and amount law ($p \propto N$) we get $p \propto \frac{N}{V} \rightarrow pV \propto N$ 2) Combining Charles' ($V \propto T$) and the pressure law ($p \propto T$) we get $pV \propto T$ 3) Combining these two proportionalities we get $pV \propto NT$ 4) Introduce the constant of proportionality $pV = NkT$, where N is the number of molecules, V is the volume in m^3, T is the temperature in K 5) An equivalent equation is $pV = nRT$, where n is the number of moles				
Boltzmann constant, k: k = 1.38 x 10^{-23}JK^{-1}. It relates temperature in kelvin to the energy of a particle Gas constant, R: k = 8.314 JmolK^{-1} $$R = N_A k$$				
Mole: The amount of substance that contains the same number of particles as there are carbons atoms in 12.0g of carbon-12 $$n = \frac{m}{M_r}$$ Where n is the number of moles, m is the mass of the sample and M$_r$ is the molar mass of the sample. [Note: Remember to double the molar mass taken from the periodic table for diatomic molecules (such as oxygen and nitrogen)]				

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Avogadro Constant: The number of particles in one mole; $N_A = 6.02 \times 10^{23} \text{mol}^{-1}$ $$N = nN_A$$ Where N is the number of molecules/atoms and n is the number of moles				
Kinetic model of gases: Matter is made of atoms/molecules in motion, we can use Newtonian mechanics to derive a theoretical explanation of the ideal gas equation. It is assumed that all collisions are perfectly elastic, so no energy is lost in collisions either with walls or other molecules				
One molecule in a box: 1) A molecule of mass, m, is moving with constant speed, c, in the x-direction. (c is used instead of v for the analysis of ideal gases) 2) Assuming all collisions are elastic, the molecule will bounce to and fro in the x-direction, with velocity changing between +c and -c as it rebounds off the right-hand wall 3) When it collides with the right-end wall there will be a change in momentum (an impulse), $\Delta p = -2mc$. With the left-hand wall $\Delta p = +2mc$. The signs show the forces are acting inwards 4) $I = F\Delta t \rightarrow F = \frac{I}{\Delta t} = \frac{2mc}{\Delta t}$ (Δt is the time between collisions <u>on the same wall</u>) 5) $S = \frac{D}{t} \rightarrow c = \frac{2x}{\Delta t} \rightarrow \Delta t = \frac{2x}{c}$ (It travels 2x to return to the same wall) 6) $F = \frac{2mc}{\Delta t} = 2mc \div \frac{2x}{c} = \frac{mc^2}{x} \rightarrow F = \frac{mc^2}{x}$ for one particle				
N molecules in a box: 1) There are N molecules in a box moving randomly 2) Approximately N/3 molecules will be travelling in each direction at a given time e.g. the number travelling in the y direction at a time t is N/3. Taking the vector velocity components in all possible directions gives the same result 3) So, the average force on the right-hand wall is $F = \frac{N}{3} \times \frac{mc^2}{x} = \frac{Nmc^2}{3x}$ 4) The pressure, $p = \frac{F}{A} = \frac{Nmc^2}{3x} \div yz = \frac{Nmc^2}{3xyz} = \frac{Nmc^2}{3V}$ 5) We arrive at the theoretical ideal gas equation $PV = \frac{1}{3}Nm\overline{c^2}$ 6) $\overline{c^2}$ is used because the molecules travel in random directions with random speeds, so an average of the speeds is required				
The mean-square speed, $\overline{c^2}$: Square c then take the average				
The root mean-square (rms) speed, $\sqrt{\overline{c^2}}$: Used to quote the molecular speeds, take the square root of the mean-square speed. Also written as c_{rms}				
1) $pV = \frac{1}{3}Nm\overline{c^2}$ 2) $pV = \frac{2}{3}N(\frac{1}{2}m\overline{c^2})$ 3) $NkT = \frac{2}{3}N(\frac{1}{2}m\overline{c^2})$ 4) $\frac{3}{2}kT = \frac{1}{2}m\overline{c^2}$ This equation relates the thermal and kinetic energy of an ideal gas				

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How the kinetic model of gases deviates for real gases: 1) Molecules interact with each other by Van der Waal's forces in pairs as stated by Newton's third. 2) As a molecule approaches the right-hand wall, the resultant force on the molecule will be to the left because there are more molecules to the left of it than the right of it 3) This decelerates the molecule, reducing the momentum the molecule strikes the wall with, decreasing the pressure experience by the wall 4) As Van der Waal's forces only act over a short range, a sample of gas with well separated molecules will display ideal gas behaviour				
Mean free path: The average distance molecules travel between collisions. Larger molecules travel less between collisions, smaller molecules travel farther between collisions.				
After many collisions, the displacement of a randomly moving molecule is hard to predict e.g. a perfume bottle is opened, the smell will slowly diffuse in all directions. Statistics can be used to analyse the movement of the molecules.				
If a molecule makes N steps in a random walk, on average it will be $\sqrt{N}$ from its starting point, hence a distance of $\sqrt{N} \times step\ length$ from its starting point. The step length is the mean free path				
Brownian Motion: The random motion of particles in a gas or liquid resulting from collisions with each other				
The time for a molecule to reach a certain distance away from its starting point is sometimes less than calculated. This is because molecules do not just move by diffusion, there are other factors such as air draughts and sudden movement				
Internal Energy, U (Units: J): The energy within a system that allows it to do work or transfer energy thermally. Internal energy is difficult to measure so instead we measure change in internal energy, ΔU				
Thermal Energy Transfer, Q (Units: J): The energy transfer associated with a temperature change e.g. frictional work can raise the temperature of a moving object				
First Law of Thermodynamics: The change in internal energy, ΔU, of a system is the sum of the work done, W, on the system and the energy transferred thermally, Q, into it $$\Delta U = W + Q$$				
W and Q can be + or -, if the system is doing work on something (e.g. pushing out a piston) then work is negative. If the system gives out thermal energy to the surrounding (e.g. during a radiative process) then Q is negative				

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$Q = mc\Delta T \ or \ E = mc\Delta\theta$ (*The latter is more commonly used in Physics*)				
Specific Heat Capacity, c (Units: Jkg^{-1}K^{-1}): The amount of energy needed to heat 1kg of substance by 1 kelvin. It is constant for a given material. The term 'specific' relates to the fixed mass of 1kg				
If the volume of the system in question does not change, no work is done, so $\Delta U = Q = E = mc\Delta\theta$				
$Q = mc\Delta\theta$ cannot be applied to a situation if the water is boiling (turning into a gas) because the temperature does not change, but work is done in breaking intermolecular bonds; so Q = 0 and hence ΔU = W				
Molar Thermal Capacity, C (Units: Jmol^{-1}K^{-1}): The value of ΔU that produces a temperature change of 1K in 1mol of the substance $$C = cM_r$$ Where c is the specific heat capacity and M$_r$ is the molar mass in kgmol^{-1}				
Mean Energy per Particle, kT: For an ideal gas, molecular interactions are negligible, so the particles have no PE. The mean energy per particle is $E = \frac{3}{2}kT$. This can be applied to other situations like vibrating atoms in a solid or photons emitted by a hot object. If an order of magnitude value for energy is needed just use $E \approx kT$				
Activation Energy, E: The energy needed to make a given process occur. Sometimes given the symbol, E_a				
Evaporation: 1) Water molecules in a sample have a range of different energies, some will have enough to escape 2) As they do the mean energy of the sample will drop 3) The molecules in the newly cooled sample will gain energy by collisions with other molecules; giving more molecules enough energy to escape and so on until the whole sample has evaporated				
A good rule of thumb is to say that processes occur at a reasonable rate if the energy required is between 15kT and 30kT. Say if an evaporation process required more than 30kT of energy, it would be very slow, and few molecules would evaporate. If it required less than 15kT, evaporation would occur at a rapid rate. – Recall kT is a good order of magnitude estimate for the energy per particle associated with a particular temperature. So this is a way of comparing the energy per particle with the energy needed for a process to occur $-\frac{E}{kT} = some \ factor \ \rightarrow \ E \ for \ process \ to \ occur = some \ factor \times kT$				

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-Consider a system where particles can gain or lose an equal sized quantum of energy, E. If the fraction of particles that acquire this energy is f, then the probability of a single particle acquiring an energy, E, is f too (Recall for a die, 1/6 of the dice will land with a 6 upwards and the chance of rolling a 6 is 1/6) -To acquire an energy 2E (to go up two energy levels) the probability is f^2				
The Boltzmann Factor, f: $$f = \frac{number\ of\ particles, N_X\ on\ X^{th}\ level}{number\ of\ particles, N_{X-1}\ on\ (X-1)^{th}\ level}$$ Where the energy difference between levels is E. -This is indicative of an exponential relationship because for each successive energy level, the number of particles will change by a constant factor (The ratio between for the number of particles on each pair of successive energy levels is constant). Recall for radioactive decay, $N = N_0 e^{-\lambda t}$, the decaying variable, N, decreases with increasing time. In the current situation, the number of particles decreases with increasing energy E (The probability gets less and less for each successive energy level), implying $N = N_0 e^{-constant \times E}$. The constant can be determined to be 1/kT, so the equation becomes; $$N = N_0 e^{-\frac{E}{kT}}$$ Where E is the activation energy (can also be written as E_a) -So, the fraction of the particles with extra energy, E, compared to the number not having that energy is the exponential term hence; $$Boltzmann\ Factor, f = e^{-\frac{E}{kT}}$$				
Boltzmann Factor and bonding: Weak intermolecular bonds can be broken by less than 10^{-20}J in some cases, so f is close to 1 even at room temperature, the process is certain to happen. Hydrogen bonds need almost ten times as much to be broken, so processes involving these bonds happen at a significant rate at room temperature. Ionic and covalent bonds require roughly 5×10^{-19}J to break, so at room temperature, f for these processes is very low				
Boltzmann Factor comparisons: <table><tr><td>Low Temperatures</td><td>High Temperatures:</td></tr><tr><td>T is low</td><td>T is high</td></tr><tr><td>$\frac{E}{kT}$ is large</td><td>$\frac{E}{kT}$ is small</td></tr><tr><td>$e^{-\frac{E}{kT}}$ is small and approaches 0</td><td>$e^{-\frac{E}{kT}}$ is large and approaches 1</td></tr></table>				

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Boltzmann Factor, viscosity and evaporation: 1)In liquids, flow occurs when molecules in the middle of the liquid gain enough energy to break the 'cage' formed by its neighbours. 2) Evaporation occurs when molecules at the surface gain enough energy to break away from the surrounding molecules to leave the liquid 3) Both of these processed occur more readily at higher temperatures, so as a liquid warms its viscosity drops and the rate of evaporation increases. The temperature dependence of viscosity and evaporation are described by the Boltzmann factor				
Statistical mechanics: The study of how the behaviour of matter and energy is governed by statistics dealing with random distributions -The total number of arrangements in a system, W, will never be decreased by a random change, in fact W generally increases when subjected to a random change. In other words, disordered systems will not, by themselves, become ordered by purely random interactions. (Leave a melted ice cube in a room, it will not by itself form into an ice cube again by itself) -This leads to the idea of entropy, S. Total entropy never decreases and generally increases. This is the Second Law of Thermodynamics. It governs the direction of chemical and physical changes. Entropy is related to the number of arrangements/ways, W, by the equation $S = k \, logW$				