Chapter 1.1 – Gases

States of Matter
- A substance may exist in only one state, or it may be present in 2 or 3 states
- **Solid** → particles are locked into a fixed position relative to others in the substance
    - Particles possess vibrational energy
    - When vibrational modes overcome attractive forces, the particles become a liquid
- **Liquid** → some attraction between particles, allowing it to take the shape of its container
- **Gas** → forces that hold particles together are completely overcome, so they escape
    - Much more significant translational energy – change in position
    - "Pressure of gas" – the collisions of gaseous particles within the walls of the container
- Boundaries of the behaviour of gases:
    - The pressure is not too high
    - The temperature is not close to the condensation temperature
- Approximations of gases that developed their boundaries:
    - Atoms/molecules of gas have no volume (negligible compared to volume of container)
    - Atoms/molecules of gas don't interact with each other

Gas Laws
- **Boyle's Law** – for a fixed amount of gas at a constant temperature, the gas volume is inversely proportional to the gas pressure
    - $P_1V_1 = P_2V_2$
- **Charles's Law** – the volume of a fixed amount of gas at a constant pressure is directly proportional to the Kelvin (absolute) temperature
    - $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- Both laws can be combined for a fixed # of moles of gas: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- **Avagadro's Law** – equal volumes of different gases, at the same temperature and pressure, contain equal numbers of molecules
    - Therefore, as the # of molecules increases, the volume of gas also increases linearly
    - $\frac{V_1}{n_1} = \frac{V_2}{n_2}$
- **Ideal Gas Law** – $PV = nRT$, where R = 8.314 L kPa mol^{-1} K^{-1} or 0.0821 L atm mol^{-1} K^{-1}

Chapter 1.2 – The Ideal Gas Law

Molar Volume of a Gas
- For solids and liquids, the mass and volume of one mole of a substance depends on the substance being considered
- Although different gases have different molar masses, Avogadro's Law states that one mole of any ideal gas always occupies the same volume at a given temp. and pressure
    - **Molar volume** – the volume of one mole of an ideal gas at STP (0^0C and 1 atm) is 22.4 L
- **Variations of the Ideal Gas Law:**
    - $d = \frac{(MM)(P)}{RT}$ or $MM = \frac{dRT}{P}$ *density* $d = \frac{M}{V}$ *mass / volume*

Dalton's Law of Partial Pressures

- In a mixture of several ideal gases, the total volume, pressure and # of moles of the mixture behave as a single gas with respect to the ideal gas law $(P_{total}V_{total} = n_{total}RT)$
- Each of the constituents in a gas mixture exhibits a **partial pressure**
- **Dalton's Law** – the total pressure exerted by a mixture of gases is the sum of all the partial pressures of these gases $(P_T = P_A + P_B + P_C)$
- **Mole Fraction**: $X_i = \frac{n_i}{n_T}$
  - The sum of all the mole fractions is always 1 $(1 = X_A + X_B + X_C)$
- At a constant temp. and volume, the pressure is directly proportional to moles:
  - $X_i = \frac{P_i}{P_T} = \frac{n_i}{n_T}$, therefore $X_i = \frac{P_i}{P_T}$
  - If gases are expressed in terms of mole fractions: $P_A = X_A P_T, P_B = X_B P_T, P_C = X_C P_T$
- At equilibrium, the rates of evaporation and condensation will be the same, and the pressure of the $H_2O_{(g)}$ will remain constant
  - Total pressure of the gas in the flask would be: $P_T = P_{air} + P_{water\ vapour}$
  - Vapour pressure of water is dependent only on the temperature of water
  - Solvents that are more volatile (have lower boiling points) have higher vapour pressures
  - **Boiling point** – the temperature at which the vapour pressure of a liquid is equal to atmospheric pressure
- **Average Molar Mass of a Mixture:**
  - $\overline{MM}$ is the weighted average of the molar masses of the different components
  - $\overline{MM} = (X_A \times MM_A) + (X_B \times MM_B) + (X_C \times MM_C)$

Gas Reaction Stoichiometry

- **Mole-volume** → reaction is given in moles and the volume of each gaseous component must be determined according to the balanced chemical equation, at the given temp. and pressure
- **Mass-volume** → mass of one component is given and the volume of a gaseous component must be determined at a stated temperature and pressure
  - Convert mass to moles, then use mole ratios
- **Volume-volume** → pressure and temperature are constant, and the stoichiometric coefficients are proportional to volume
- **Pressure-pressure** → pressure is directly proportional to the mole amount at constant volume and temperature

The Kinetic Molecular Theory of Gases

- **Assumptions of the theory:**
  - The size of the molecules is negligible when compared to the space between them and the volume of container
  - Molecules moving through space are considered to be point particles (no vibrational or rotational energy, just translational – kinetic energy)
  - The collisions between molecules and between molecules & the container, are considered to be elastic and at a constant velocity, in a straight line
  - No attractive or repulsive forces between molecules
- Average kinetic energy of the molecules is proportional to the temperature
  - $\overline{E_k} = \frac{1}{2}m\overline{u^2}$, where m is the mass of one molecule and $\overline{u^2}$ is the mean-square speed
  - Not all the molecules are moving at the same speed at the same time

- 3 factors determine the pressure of a gas in a container: mass of each molecule, speed of the molecules, and the frequency of collisions (which is related to the # of molecules, *N* and the volume of the container)
- Basic equation of the kinetic theory of gases: $P = \frac{1}{3}\frac{N}{V}m\overline{u^2}$, which leads to $\overline{u^2} = \frac{3RT}{MM}$
- At any given moment, some molecules are moving faster and some are moving slower
- **Root-Mean-Square Speed of Gas Molecules**: $u_{rms} = \sqrt{\frac{3RT}{MM}}$, where molar mass is in kg
- The u_{rms} for 2 gases is inversely proportional to the square root ratio of their molar masses:
  - $\frac{u_{rms(A)}}{u_{rms(B)}} = \sqrt{\frac{MM_B}{MM_A}}$

涛漏物
Effusion and Graham's Law
- **Effusion** – flow of gas molecules at low pressures through tiny pores or pinholes
- The rate of effusion of a gas is inversely proportional to the square root of its molar mass
  - Therefore lighter molecules effuse more rapidly than heavier ones
- When 2 gases effuse through a small hole, Graham's Law can be written as:
  - $\frac{rate\ of\ effusion\ of\ A}{rate\ of\ effusion\ of\ B} = \frac{N_A}{N_B}\sqrt{\frac{MM_B}{MM_A}}$
- The gas that emerges from a container is *enriched* 浓缩 in the lighter component
- The $\sqrt{\frac{MM_B}{MM_A}}$ proportion of Graham's Law is called the enrichment factor, where the heavier MM is the numerator of the fraction
- If a mixture of 2 gases has the same # of moles of each gas, after the mixture effusing for a certain length of time, the mixture that effused would have a proportionally higher ratio of the lighter gas

Chapter 2.1 – Heat, Work and Energy

Thermodynamics
- The study of the energy changes involved in physical and chemical processes
  - *Physical process* → heating and cooling of substances, phase changes, and V/P changes
  - *Chemical process* → chemical reactions
- **Thermochemistry** – investigates the heat flow that occurs during these reactions
  - Rates of most chemical reactions increases as temperature increases

Systems and Surroundings
- **Heat flow** – the transfer of thermal energy in the process being studied
  - Heat always flows from the warmer object to the cooler object
- **System** – the part of the universe that one wants to study
- **Surroundings** – the remaining parts of the universe that can interact (e.g. exchange energy) with a system under study
- In a chemical reaction, the system is the reactants and the surroundings is the vessel (flask) in which the reaction takes place, including the air and any other material making thermal contact with the vessel, such as the solution in which the reaction occurs