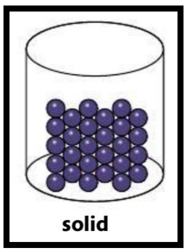
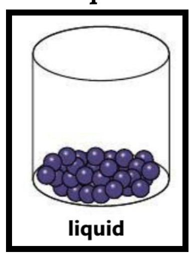
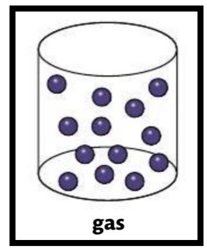
1.1 Introduction to the particulate nature of matter and chemical change

Elements are made of only one type of atom and are the building blocks for the formation of compounds. ***Compounds*** form as a result of a chemical reaction between fixed ratios of atoms of 2 or more different elements. The physical and chemical properties of elements and their compounds are different.

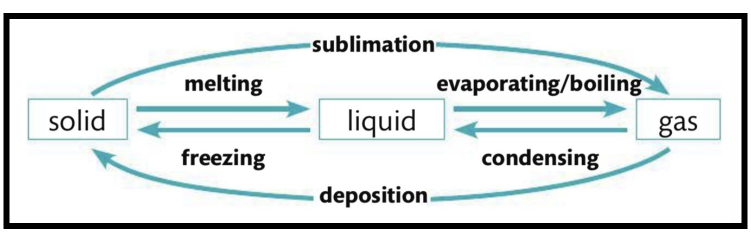
Mixtures are substances that contain more than one element and/or compound that are not chemically bonded together. This means that the constituents retain their individual properties. There are 2 types of mixtures:

- ➔ ***Homogenous mixtures*** - uniform composition and properties throughout. Eg: air
- ➔ ***Heterogeneous mixtures*** - non-uniform compositions, so properties aren't the same throughout. Eg: water, oil

Matter exists in different ***states*** depending on the temperature and pressure:

Solid	**Liquid**	**Gas**
		
Increasing temperature Increasing kinetic energy of particles		
• Particles closely packed • Strong attractive forces • Particles vibrate • Fixed shape and volume	• Particles loosely packed • Weak attractive forces • Particles slide past • Fixed volume, not shape	• Particles spread out • Negligible IM forces • Particles move at high speed • No fixed volume or shape

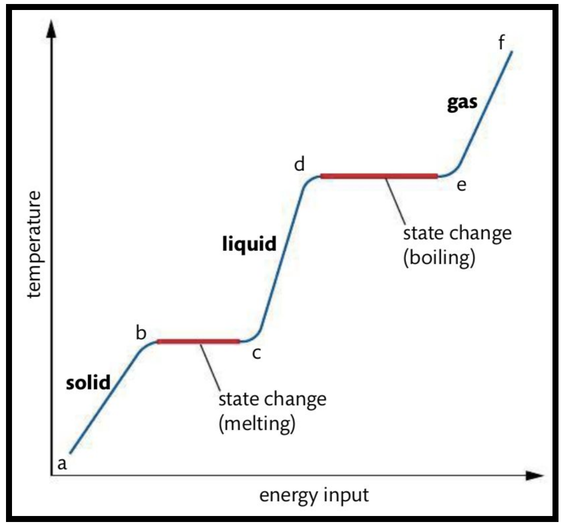
Liquids and gases are **fluids**. This means that they can flow. This allows **diffusion** to occur. Diffusion is the process by which the particles of a substance become evenly distributed, as a result of their random movements.



State changes. Evaporation occurs at all temp. up to bp and at the surface. Boiling is only at bp and throughout the liquid. BP is determined by when vapor pressure reaches an external pressure.

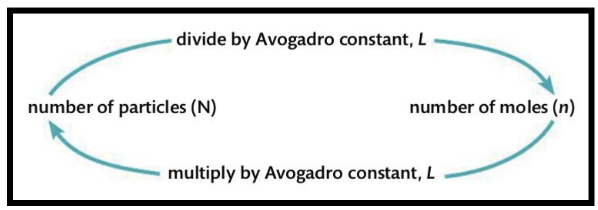
The Heating Curve

- a→b - particles gain KE, so the temperature of the solid increases (increased average KE)
- b→c (**melting point**) - To change from solid to liquid state, particles are pulled further apart. In doing so, PE increases (rather than KE) so temperature remains constant
- c→d - particles gain KE, so the temperature of the liquid increases (increased average KE)
- D→e (**boiling point**) - similar to melting, converting liquid to gas requires an increase in PE and not KE, so temperature is constant
- e→f - particles gain KE, so the temperature of the gas increases (increased average KE)



1.2 The mole concept

Mole - the amount of a substance that contains the same number of units as the number of carbon atoms in 12g of carbon-12. There are 6.02 x 10^{23} atoms of carbon in 12g of carbon-12. This value is called ***Avogadro's constant***.



Relative atomic mass - the relative atomic mass, A_r, for an element is the average mass of its naturally-occurring isotopes relative to the mass of a carbon-12 atom (mass of exactly 12 units)

Relative molar mass - the relative molecular mass, M_r, is the sum of the A_r in a molecule

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} \qquad n\,(\text{mol}) = \frac{m\,(\text{g})}{M\,(\text{g mol}^{-1})}$$

Empirical formulas give us the simplest ratio of elements in a compound

Suppose percentages are given, assume there were 100g of that compound. Divide each mass by the A_r of the respective element to find a ratio for the number of atoms of each element in the empirical formula.

For eg: A compound consists of 80% carbon and 20% H. Find its empirical formula:

(80/100) x 100g = 80g of C (20/100) x 100g = 20g of H

C = 80/12 = 6.67

H = 20/1 = 20

The ratio of combination ⟶ 6.67:20 ⟶ 1:3, ∴ empirical formula = $\mathbf{CH_3}$

To go backward, take the total atomic mass of each atom as a percentage of the molar mass (M).

Molecular formulae show the actual number of atoms in a compound.

Will be given an empirical formula, or a continuation where you have already found it. Using the empirical formula and a given M_r, the molecular formula can be found:

- Find a constant 'n' which = M_r / empirical mass
- Multiply each of the numbers in the empirical formula with 'n'

Eg: The M_r of a compound with the empirical formula HO is 34. What is its molecular formula?

n = $34 gmol^{-1}$/ (1x1) + (1x16) = 34/17 = 2
∴ Molecular formula = H_2O_2

1.3 Reacting masses and volumes

Limiting reactants, reactants in excess, and theoretical yield

The reactant that determines the quantity of product is known as the *limiting reactant*. Other reactants are therefore unused, and so are in *excess*. By identifying the limiting reactant, we can calculate the *theoretical or expected yield*.

Experimental yield and % yield

In the calculation of theoretical yield, we assume that there was no loss, impurities present, wastage, or incomplete reaction. In reality, most of the above take place in chemical reactions, so the **experimental yield** differs from the theoretical yield. Using these two yields, the % yield is calculated:

$$\boxed{\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\%}$$

Avogadro's Law

Equal volumes of gases, when measured at the same temp. and pressure, contain an equal no. of particles

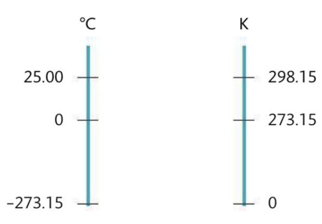
$$\boxed{\text{number of moles of gas } (n) = \frac{\text{volume } (V)}{\text{molar volume}}}$$

Based on this law, one mole of any gas must have the same volume given that temperature and pressure are constant. At standard temperature (0°C) and pressure (100kPa), one mole of gas has a volume of 22.7 dm^3/mol. At room temp. (25°C) this increases to 24 dm^3/mol

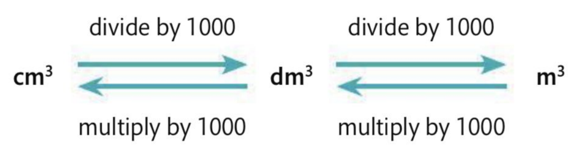
Units

Temperature

temperature (K) = temperature (°C) + 273.15



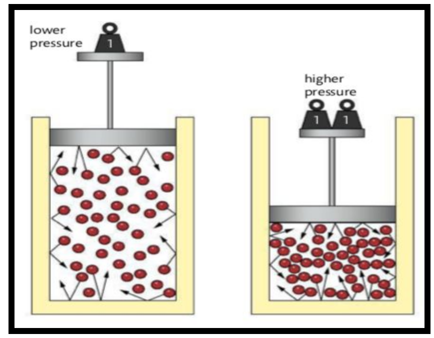
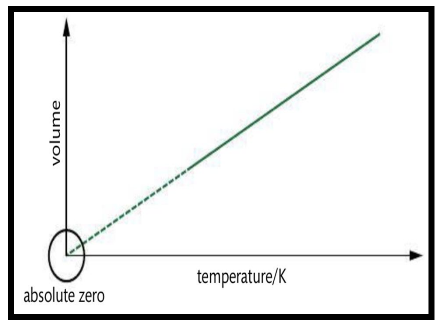
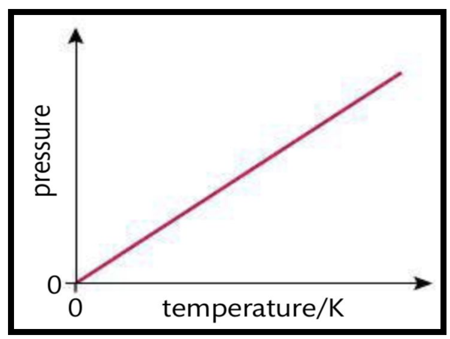
Volume



Pressure

Pa = N/m^2

Gas Laws

The relationship between volume and pressure Volume of a gas is always the volume of its container. Pressure is caused by particles of gas colliding with the walls of the container. If the temperature is constant, increasing pressure on a fixed mass of gas results in a reduction in volume. $\boxed{\begin{array}{c} P \propto \dfrac{1}{V} \\ PV = \text{a constant} \end{array}}$	
The relationship between volume and temperature An increase in temperature results in particles containing more KE, which allows them to move around faster. Assuming pressure is constant, this results in an increased volume, as particles need more space to move in. **Note: if the temperature units are in °C, then the graph starts at -273** $\boxed{\begin{array}{c} V \propto T \\ \dfrac{V}{T} = \text{a constant} \end{array}}$	
The relationship between pressure and temperature An increase in temperature results in particles having greater KE. This means they move faster, increasing the frequency and size of collisions with the walls of the container. This results in increased pressure. $\boxed{\begin{array}{c} P \propto T \\ \dfrac{P}{T} = \text{a constant} \end{array}}$	

The three laws that can be applied to a fixed mass of gas can be summarized as follows:	These can be combined to give an equation for a fixed mass of gas:
$\boxed{\begin{array}{c} P \propto \dfrac{1}{V} \text{ at constant temperature} \\ V \propto T \text{ at constant pressure} \\ P \propto T \text{ at constant volume} \end{array}}$	$\boxed{\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}}$ Where 1 and 2 refer to initial and final conditions respectively.

Derivation of the ideal gas equation from the combined gas equation and Avogadro's law

The combined gas equation tells us that:

$$\boxed{\frac{PV}{T} = \text{a constant}}$$

The value of the constant is directly proportional to the number of moles, n. So:

$$\boxed{\frac{PV}{T} \propto n}$$

This can be made into an equation by introducing a constant, **R**, known as the **universal gas constant**:

$\boxed{\frac{PV}{T} = nR}$, which is commonly written as **PV = nRT**

R can be calculated by substituting in values for other quantities. It should equate to **8.31 J/K mol**.

For all calculations, the SI units **must** be used. To avoid mistakes, follow the rules below:

- Pressure, P: must be in **Pa**; if kPa is given, multiply by 10^3
- Volume, V: must be **m^3**; if dm^3 are given, divide 10^3; if cm^3 are given, divide by 10^6
- Number of moles, n: this is given by dividing mass over M_r of the substance
- Temperature, T: must be in **K**; if °C is given, add 273.15

Non-ideal behavior of gases

This occurs if:

- **Pressure is high**
  - At high pressures, gas molecules get more crowded and the amount of space between the molecules is reduced. The volume we use in PV = nRT is the empty volume in the container - at low pressures, this volume can be assumed to be the volume of the container (as gas particles don't take up that much space). However, at high pressures, the volume of the gas particles themselves becomes significant, so real gases take up a larger volume than predicted by the ideal gas law.
- **Temperature is low**
  - At high temperatures, particles have enough KE to allow us to deem intermolecular forces as negligible. However, at low temperatures, particles have low KE, which is not enough to overcome IM forces. If pressure is to remain constant, the volume decreases.

Concentration

To find the concentration of a substance, we use the following equation:

$$\text{concentration of solution (mol dm}^{-3}\text{)} = \frac{\text{amount of solute (mol)}}{\text{volume of solution (dm}^3\text{)}} \quad \text{or} \quad c = \frac{n}{V}$$

Concentration can also be measured in **parts per million (ppm).** It denotes one part per 10^6 by mass.

$$\text{parts per million (ppm)} = \frac{\text{mass of component}}{\text{total mass of solution}} \times 10^6$$

Dilute solutions

As a solution is diluted, the number of moles remains the same, but as they spread over a larger volume, the concentration decreases.

$$c_1V_1 = c_2V_2$$

Determining the concentration of a substance

In most cases, **titration** is carried out to find the concentration. In the example of finding the concentration of HCl, a **pipette** is used to measure a given volume of the solution with an unknown concentration, and placed into a **conical flask**. Next, NaOH, a base, is placed in a **burette** above the conical flask. The point at which the two solutions have reacted completely, the **equivalence point**, is usually determined by an **indicator** that is added to the solution in the conical flask and changes color at its **endpoint**.

Back titration

- First, write the equation for the reaction
- Look for the reactant whose vol. and conc. are given and calculate its no. of moles from n = cV
- Use this with the mole ratio in the equation to determine the no. of moles of the other reactant
- Use the no. of moles and vol. of the second reactant to calculate its conc. from c = n/V