

KWV EDUCATION



KWV PHYSICS

: WINTER BOOTCAMP

GRADE

:11

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QUANTITATIVE ASPECTS OF CHEMICAL CHANGE

The Mole

Atoms, molecules and ions are too small to count, and there are so many particles in even the smallest sample of a substance.

There are more particles of water in a teaspoon than there are teaspoons of water in all the oceans.

Rather than dealing with the particles individually, we deal with a special number of particles.

The **mole** is a name for a special number. Many numbers have names, such as:

2 = pair

3 = hat-trick

12 = dozen

A **mole** of particles is an amount of $6,02 \times 10^{23}$ particles. $6,02 \times 10^{23}$ is known as **Avogadro's number**, N_A . Avogadro's number (N_A) is too big to imagine.

602 000 000 000 000 000 000 000 .

This many grains of sand, piled on the surface of the earth would almost reach the moon.

The mole is defined as the number of particles or atoms in 12,0 g of Carbon -12.

Molar Mass

Particles are too small to weigh individually.

Molar mass (M) is defined as the mass of one mole of particles (atoms, molecules or formula units) and is measured in the unit $\text{g} \cdot \text{mol}^{-1}$.

$$\text{number of mole (mol)} \leftarrow n = \frac{m}{M} \rightarrow \begin{array}{l} \text{mass of substance (g)} \\ \text{molar mass (g} \cdot \text{mol}^{-1}) \end{array}$$

Relative atomic mass (A_r) is the average mass of an atom compared to the mass of a Carbon 12 atom. It is measured in atomic mass units (**amu**).

Molar mass (M) of an element is equal to the magnitude of **relative atomic mass (A_r)** in amu. This is found on the periodic table. See the table below for other substances:

Type of substance	Particles	Example	Formula	Molar mass $\text{g} \cdot \text{mol}^{-1}$
Element	Atoms	Neon	Ne	20
Covalent compound	Usually Molecules	Carbon dioxide	CO_2	$12+32$ $= 44$
Ionic compound	Ions (formula units)	Salt	NaCl	$23+35,5$ $= 58,5$
Metallic compound	Positive kernels and delocalized electrons	Gold	Au	197

EXAMPLE:

What is the relative formula mass of Calcium sulphate (CaSO_4)?

$$M_R(\text{CaSO}_4) = A_R(\text{Ca}) + A_R(\text{S}) + (4 \times A_R(\text{O}))$$

$$= 40 + 32 + (4 \times 16)$$

$$= 136 \text{ (no unit)}$$

EXAMPLE:

What is the molar mass of

Calcium sulphate (CaSO_4)?

$$136 \text{ g} \cdot \text{mol}^{-1}$$

EXAMPLE:

What is the relative molecular mass of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$)?

$$M_R(\text{C}_{12}\text{H}_{22}\text{O}_{11}) = 12 \times A_R(\text{C}) + 22 \times A_R(\text{H}) + 11 \times A_R(\text{O})$$

$$= (12 \times 12) + (22 \times 1) + (11 \times 16)$$

$$= 342 \text{ (no unit)}$$

EXAMPLE:

What is the molar mass of

sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) ?

$$342 \text{ g} \cdot \text{mol}^{-1}$$

Percentage Composition

$$\text{Percentage composition of element} = \frac{\text{molar mass of element}}{M_R \text{ of compound}} \times 100$$

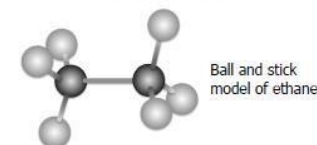
Consider these iron ores: haematite and magnetite – which contains more iron by mass?

Ore	Haematite	Magnetite
Formula	Fe_2O_3	Fe_3O_4
Relative molecular mass	$(2 \times 56) + (3 \times 16)$ $= 160$	$(3 \times 56) + (4 \times 16)$ $= 232$
% iron by mass	$[(2 \times 56) / 160] \times 100$ $= 70\%$	$[(3 \times 56) / 232] \times 100$ $= 72\%$

$\therefore$ magnetite contains more iron

Different types of Chemical Formulae

Consider the substance ethane



It also can be represented using a formula. There are three types of formulas we use:

Molecular formula	Actual number of each atom.	Eg. C_2H_6
Empirical formula	Simplest whole number ratio of the atoms.	Eg. $\text{C}_2\text{H}_6 \rightarrow \text{CH}_3$
Structural formula	Shows how the atoms are joined.	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$

QUANTITATIVE ASPECTS OF CHEMICAL CHANGE

Concentrations of solutions

Solutions are homogeneous (uniform) mixtures of two or more substances. A solution is formed when a solute dissolves in a solvent. The solvent and solute can be a gas, liquid, or solid. The most common **solvent** is liquid water. This is called an **aqueous solution**.

Solution	Solute	Solvent
salt water	Salt	water
Soda water	Carbon dioxide	water

Concentration

The **concentration** of a solution is the number of mole of solute per unit volume of solution.

$$\text{concentration (mol} \cdot \text{dm}^{-3}) \leftarrow c = \frac{n}{V} \rightarrow \text{volume (dm}^3\text{)}$$

number of moles (mol)

can also be calculated with

$$c = \frac{m}{MV}$$

i.e. The number moles of solute per 1 dm³ of solution i.e. mol·dm⁻³. If a solution of **potassium permanganate KMnO₄** has a concentration of 2 mol·dm⁻³ it means that for every 1 dm³ of solution, there are 2 moles of KMnO₄ dissolved in the solvent.

Note:

1000 cm³ = 1 dm³ (= 1 litre)
1 cm³ = 0,001 dm³
1 cm³ = 1 ml

EXAMPLE:

A solution contains 10 g of sodium hydroxide, NaOH, in 200 cm³ of solution. Calculate the concentration of the solution.

$$\begin{aligned} n(\text{NaOH}) &= \frac{m}{M} \\ &= \frac{10}{23 + 16 + 1} \\ &= 0,25 \text{ mol} \end{aligned}$$

$$V = 200 \text{ cm}^3 = 0,2 \text{ dm}^3$$

$$\begin{aligned} c(\text{NaOH}) &= \frac{n}{V} \\ &= \frac{0,25}{0,2} \\ &= 1,25 \text{ mol} \cdot \text{dm}^{-3} \end{aligned}$$

EXAMPLE:

Calculate the mass of solute in 600 cm³ of 1,5 mol·dm⁻³ sodium chloride solution.

$$V = 600 \text{ cm}^3 = 0,6 \text{ dm}^3$$

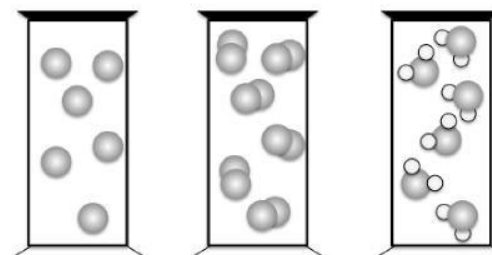
$$\begin{aligned} M(\text{NaCl}) &= 23 + 35,5 \\ &= 58,5 \text{ g} \cdot \text{mol}^{-1} \end{aligned}$$

$$\begin{aligned} n &= cV \\ &= 1,5 \times 0,6 \\ &= 0,9 \text{ mol} \end{aligned}$$

$$\begin{aligned} m &= nM \\ &= 0,9 \times 58,5 \\ &= 52,65 \text{ g} \end{aligned}$$

Molar Volumes of Gases

If different gases have the same volume under the same conditions of temperature and pressure, they will have the same number of molecules.

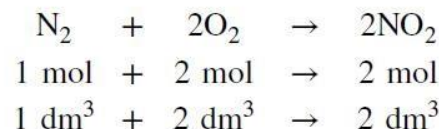


The **molar volume** of a gas, V_M , is the volume occupied by one mole of the gas.

V_M for **all gases** at STP is 22,4 dm³·mol⁻¹.

Standard Temperature and Pressure (STP) is 273 K (0°C) and 1,01x10⁵ Pa.

This also means that for reactions at constant temperature and pressure, gas volumes will react in the same ratio as the molar ratio.



$$\text{number of moles (mol)} \leftarrow n = \frac{V}{V_M} \rightarrow \text{volume of gaseous substance (dm}^3\text{)}$$

molar gas volume at STP (22,4 dm³ · mol⁻¹)

EXAMPLE:

A gas jar with a volume of 224 cm³ is full of **chlorine gas**, at STP. How many **moles** of chlorine gas are there in the gas jar?

$$\begin{aligned} n &= \frac{V}{V_M} \\ &= \frac{0,224}{22,4} \\ &= 0,01 \text{ mol} \end{aligned}$$

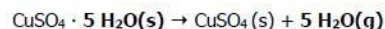
QUANTITATIVE ASPECTS OF CHEMICAL CHANGE

Water of crystallization

Some ionic crystals trap a certain number of water molecules between the ions as they are forming. These water molecules are known as "Water of crystallization".

Eg. Hydrated copper sulphate:

$\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ has 5 water molecules per formula unit. When the hydrated salt crystals are heated, the water molecules evaporate off, leaving the anhydrous salt behind.



To calculate the number of moles of water of crystallization:

1. Calculate the mass of water that evaporated off.
2. Calculate the moles of water.
3. Calculate the moles of anhydrous salt.
4. Determine the ratio of water to anhydrous salt.
5. Write the formula for the hydrated salt.

Note:

The dot in the formula ($\cdot$) between the salt and the water means that a light bond is formed. It is **NOT** a multiplication dot.

EXAMPLE:

13,2 g of a sample of zinc sulphate, $\text{ZnSO}_4 \cdot x\text{H}_2\text{O}$, was heated in a crucible. Calculate the number of moles of water of crystallisation if 7,4 g of solid remained.

$$1. m(\text{H}_2\text{O}) = 13,2 \text{ g} - 7,4 \text{ g} \\ = 5,8 \text{ g}$$

$$2. n(\text{H}_2\text{O}) = \frac{m}{M} \\ = \frac{5,8}{18} \\ = 0,32 \text{ mol}$$

$$3. n(\text{ZnSO}_4) = \frac{m}{M} \\ = \frac{7,4}{161} \\ = 0,046 \text{ mol}$$

$$4. \text{ratio} = \frac{\text{mol water}}{\text{mol anhydrous salt}} \\ = \frac{0,32}{0,046} \\ = 1 : 7$$

$$5. \therefore \text{formula} = \text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$$

Calculating the Empirical Formula of a Compound from Mass

Empirical formula is the chemical formula of a compound that shows the smallest whole number ratio of the atoms.

To calculate the empirical formula of a compound from mass:

1. Determine the mass of the elements.
2. Determine mol of each substance.
3. Simplify the atomic ratio.

EXAMPLE:

In a combustion reaction 0,48 g of magnesium ribbon is burnt. The amount of magnesium oxide produced is 0,80 g. Calculate the empirical formula for magnesium oxide.

Steps	Magnesium	Oxygen
Step 1: Mass of element	0,48 g	0,80 – 0,48 = 0,32 g
Step 2: Mol (divide by mass of 1 mol)	$n = m / M$ $= 0,48 / 24$ $= 0,02 \text{ mol}$	$n = m / M$ $= 0,32 / 16$ $= 0,02 \text{ mol}$
Step 3: Atom ratio (divide by smallest no in ratio)	1	1

Empirical formula: MgO

EXAMPLE:

A sample of an oxide of copper contains 8 g of copper combined with 1 g of oxygen. Find the empirical formula of the compound.

Steps	Copper	Oxygen
Step 1: Mass of element	8 g	1 g
Step 2: Mol (divide by mass of 1 mol)	$n = m / M$ $= 8 / 63,5$ $= 0,126 \text{ mol}$	$n = m / M$ $= 1 / 16$ $= 0,0625 \text{ mol}$
Step 3: Atom ratio (divide by smallest no in ratio)	$0,126 / 0,0625$ ≈ 2	$0,0625 / 0,0625$ $= 1$

Empirical formula: Cu_2O

Calculating the Empirical Formula from Percentage Composition

The empirical formula of a compound can also be found from its percentage composition. We assume that 100 g of the compound is analysed, then each percentage gives the mass of the element in grams in 100 g of the compound.

An oxide of sulphur contains 40% sulphur and 60% oxygen by mass. Determine the empirical formula of this oxide of sulphur.

Steps	Sulphur	Oxygen
Step 1: % of element	40	60
Step 2: Mass of element (g)	40	60
Step 3: Mol	$n = m / M$ $= 40 / 32$ $= 1,25 \text{ mol}$	$n = m / M$ $= 60 / 16$ $= 3,75 \text{ mol}$
Step 4: Smallest mol ratio	$1,25 / 1,25$ $= 1$	$3,75 / 1,25$ $= 3$

Empirical formula: SO_3

QUANTITATIVE ASPECTS- REACTIONS

Empirical formula to Molecular Formula

The **empirical formula** is the simplest whole number ratio of atoms in a molecule. The **molecular formula** is the actual ratio of the atoms in a molecule.

The molecular formula can be calculated from the empirical formula and the relative molecular mass.

STEPS TO DETERMINE MOLECULAR FORMULA:

1. Determine the empirical formula (if not given).
2. Determine the molar mass of the empirical formula.
3. Determine the ratio between molecular formula and empirical formula.
4. Multiply the ratio into the empirical formula

EXAMPLE:

Butene has the empirical formula CH_2 . The molecular mass of butene is $56 \text{ g} \cdot \text{mol}^{-1}$. Determine the molecular formula of butene.

1. Empirical formula given: CH_2

$$\begin{aligned} 2. \quad M(\text{CH}_2) &= 12 + 1 + 1 \\ &= 14 \text{ g} \cdot \text{mol}^{-1} \end{aligned}$$

$$\begin{aligned} 3. \quad \text{ratio number} &= \frac{\text{molecular formula mass}}{\text{empirical formula mass}} \\ &= \frac{56}{14} \\ &= 4 \end{aligned}$$

$$4. \quad \text{CH}_2 \times 4 = \text{C}_4\text{H}_8$$

Balanced chemical equations



The balanced chemical equation is the "recipe". It tells us:

- The reactants used i.e. Nitrogen and Hydrogen
- The products produced. i.e. ammonia
- The states of the reactants and products (all gases here)
- The molar ratio of the reactants to products.

1 molecule of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia.

1 mole of nitrogen reacts with 3 moles of hydrogen to form 2 moles of ammonia.

Approach to reaction stoichiometry

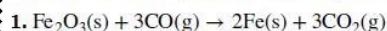
1. Write a balanced chemical equation.
2. Change the 'given' amount into mole (use limiting reactant if applicable).
3. Determine the number of mole of the 'asked' substance using the mole ratio.
4. Determine the 'asked' amount from the number of mole.

EXAMPLE:

To obtain iron, iron (III) oxide reacts with carbon monoxide according to the following word equation:

iron (III) oxide + carbon monoxide $\rightarrow$ iron + carbon dioxide

What mass of iron is produced from 48 g of iron (III) oxide?



$$\begin{aligned} 2. \quad n(\text{Fe}_2\text{O}_3) &= \frac{m}{M} \\ &= \frac{48}{160} \\ &= 0,3 \text{ mol} \end{aligned}$$

$$\begin{aligned} 3. \quad \text{Fe}_2\text{O}_3 &: \quad \text{Fe} \\ 1 &: \quad 2 \\ 0,3 \text{ mol} &: \quad 0,6 \text{ mol} \end{aligned}$$

$$\begin{aligned} 4. \quad m(\text{Fe}) &= nM \\ &= (0,6)(56) \\ &= 33,6 \text{ g} \end{aligned}$$

GIVEN VALUE

GIVEN MOL

ASKED MOL

ASKED VALUE

STEP TWO:

$$\begin{array}{|c|c|c|} \hline \text{Solid:} & \text{Solution:} & \text{Gas:} \\ \hline n = \frac{m}{M} & n = cV & n = \frac{V}{22,4} \\ \hline & & (\text{At STP}) \end{array}$$

STEP THREE: Use molar ratios

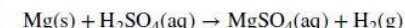
STEP FOUR:

$$\begin{array}{|c|c|c|} \hline \text{Solid:} & \text{Solution:} & \text{Gas:} \\ \hline m = nM & c = \frac{n}{V} & V = n \times 22,4 \\ \hline & & (\text{At STP}) \end{array}$$

Limiting Reactants

In a reaction between two substances, one reactant is likely to be used up completely before the other. This limits the amount of product formed.

Consider the reaction between magnesium and dilute sulphuric acid. The balanced chemical equation is



This means that 1 mole of magnesium reacts with 1 mole of sulphuric acid. Both reactants will be completely used up by the time the reaction stops.

What happens if 1 mole of magnesium reacts with 2 mole of sulphuric acid? There is now insufficient magnesium to react with all of the sulphuric acid. 1 mole of sulphuric acid is left after the reaction.

All of the magnesium is used up, We say the magnesium is the **limiting reactant**. Some sulphuric acid is left after the reaction. We say the sulphuric acid is in **excess**.

The amount of limiting reactant will determine:

- The amount of product formed.
- The amount of other (excess) reactants used.

Determining limiting reactants

1. Calculate the number of moles of each reactant.
2. Determine the ratio between reactants.
3. Determine limiting reactant using the ratios.

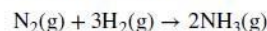
NOTE:

If one reactant is in **excess**, it means that there is more than enough of it. If there are only 2 reactants and one is in excess, it means that the other is the limiting reactant.

QUANTITATIVE ASPECTS OF CHEMICAL CHANGE

EXAMPLE:

A 8,4 g sample of nitrogen reacts with 1,5 g hydrogen. The balanced equation is:



Determine (a) which reactant is the limiting reactant, and (b) the mass of ammonia that can be produced.

(a)**1.**

$$\begin{aligned} n(\text{N}_2) &= \frac{m}{M} \\ &= \frac{8,4}{28} \\ &= 0,3 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{H}_2) &= \frac{m}{M} \\ &= \frac{1,5}{2} \\ &= 0,75 \text{ mol} \end{aligned}$$

2.

$$\begin{array}{lcl} \text{N}_2 & : & \text{H}_2 \\ 1 & : & 3 \\ \hline 0,3 \text{ mol} & : & 0,9 \text{ mol} \\ 0,25 \text{ mol} & : & 0,75 \text{ mol} \end{array}$$

3.

If all nitrogen is used, 0,9 mol hydrogen is needed. However, only 0,75 mol hydrogen is available. The hydrogen will run out first, therefore hydrogen is the limiting reactant.

(b)

Because the hydrogen is the limiting reactant, it will determine the mass of ammonia produced:

$$\begin{array}{lcl} \text{H}_2 & : & \text{NH}_3 \\ 3 & : & 2 \\ \hline 0,75 \text{ mol} & : & 0,5 \text{ mol} \end{array}$$

$$\begin{aligned} m(\text{NH}_3) &= nM \\ &= (0,5)(17) \\ &= 8,5 \text{ g} \end{aligned}$$

Percentage Purity

Sometimes chemicals are not pure and one needs to calculate the percentage purity. Only the pure component of the substance will react. For an impure sample of a substance:

$$\text{Percentage purity} = \frac{\text{Mass of pure substance}}{\text{Mass of impure substance}} \times 100$$

STEPS TO DETERMINE THE PERCENTAGE PURITY

1. Determine moles of products.
2. From the balanced formula, determine the ratio between reactants and products.
3. Using the ratio, determine the number of moles of reactants.
4. Determine the mass of pure reactant.
5. Calculate the percentage purity of the sample.

EXAMPLE:

An impure sample of calcium carbonate, CaCO_3 , contains calcium sulphate, CaSO_4 , as an impurity. When excess hydrochloric acid was added to 6 g of the sample, 1200 cm^3 of gas was produced (measured at STP). Calculate the percentage purity of the calcium carbonate sample. The equation for the reaction is:



$$1. n(\text{CO}_2) = \frac{V}{V_M} = \frac{1,2}{22,4} = 0,054 \text{ mol}$$

2.

$$\begin{array}{lcl} \text{CaCO}_3 & : & \text{CO}_2 \\ 1 & : & 1 \end{array}$$

3.

$$\begin{array}{lcl} \text{CaCO}_3 & : & \text{CO}_2 \\ 1 & : & 1 \\ \hline 0,054 & : & 0,054 \end{array}$$

$\therefore 0,054 \text{ mol CaCO}_3$ reacted

$$4. m(\text{CaCO}_3) = nM = (0,054)(40 + 12 + 16 + 16 + 16) = 5,4 \text{ g}$$

5.

$$\text{Percentage purity} = \frac{\text{Mass of pure substance}}{\text{Mass of impure substance}} \times 100$$

$$\text{Percentage purity} = \frac{5,4}{6,0} \times 100$$

$$\text{Percentage purity} = 90 \%$$

Percentage Yield

When you make a chemical in a laboratory a little of the chemical is always lost, due to evaporation into the surrounding air, or due to a little being left in solution. Some of the reactants may not react. We say that the reaction has not run to completion.

This results in the amount of the chemical produced always being less than the maximum theoretical amount you would expect. We can express this by the percentage yield:

$$\text{Percentage yield} = \frac{\text{Mass of product produced}}{\text{Maximum theoretical mass of product}} \times 100$$

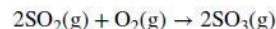
Percentage yield is usually determined using mass, but can also be determined with mol and volume.

STEPS TO DETERMINE THE PERCENTAGE YIELD

1. Determine moles of reactants.
2. From the balanced formula, determine the ratio between reactants and products.
3. Using the ratio, determine the number of moles of products.
4. Determine the theoretical mass of product.
5. Calculate the percentage yield.

EXAMPLE:

128g of sulphur dioxide, SO_2 , was reacted with oxygen to produce sulphur trioxide, SO_3 . The equation for the reaction is:



140 g of SO_3 was produced in the reaction. Calculate the percentage yield of the reaction.

$$1. n(\text{SO}_2) = \frac{m}{M} = \frac{128}{64} = 2 \text{ mol}$$

2.

$$\begin{array}{lcl} \text{SO}_2 & : & \text{SO}_3 \\ 2 & : & 2 \\ \hline 1 & : & 1 \end{array}$$

3.

$$\begin{array}{lcl} \text{SO}_2 & : & \text{SO}_3 \\ 1 & : & 1 \\ \hline 2 & : & 2 \\ \hline \therefore 2 \text{ mol SO}_3 \end{array}$$

$$4. m(\text{SO}_3) = nM = (2)(32 + 16 + 16 + 16) = 160 \text{ g}$$

5.

$$\text{Percentage yield} = \frac{\text{Mass of product produced}}{\text{Maximum theoretical mass of product}} \times 100$$

$$\text{Percentage yield} = \frac{140}{160} \times 100$$

$$\text{Percentage yield} = 87,5 \%$$

ENERGY CHANGE

The energy change that takes place occurs because of bonds being broken and new bonds being formed.

When bonds are broken, energy is absorbed from the environment.

When bonds are formed, energy is released into the environment.

The net energy change will determine if the reaction is endothermic or exothermic.

ACTIVATION ENERGY

In order to start a reaction, energy first needs to be absorbed to break the bonds. This energy is known as the **activation energy**- the minimum energy required to start a chemical reaction.

Once the bonds have been broken, the atoms in the chemical system form an **activated complex**- a temporary transition state between the reactants and the products.

ENTHALPY AND ENTHALPY CHANGE

Enthalpy (H) is the total amount of stored chemical energy (potential energy) of the reactants and the products. During chemical reactions, energy can be exchanged between the chemical system and the environment, resulting in a change in enthalpy. This change in enthalpy, ΔH , represents the heat of the reaction measured in $\text{kJ}\cdot\text{mol}^{-1}$.

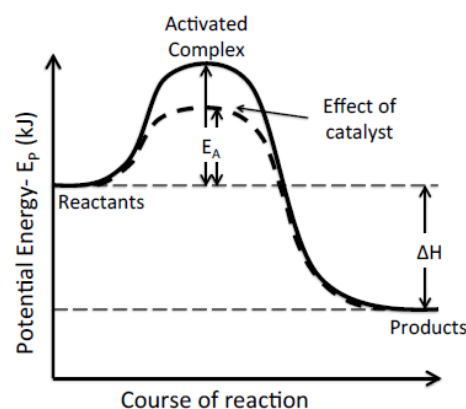
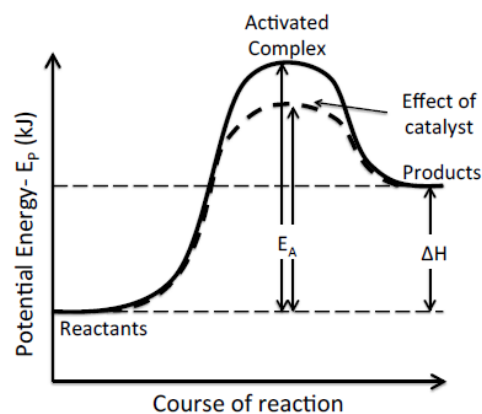
The heat of reaction (ΔH) is the net change of chemical potential energy of the system.

CHEMICAL SYSTEM AND THE ENVIRONMENT

The chemical system is the reactant and product molecules.

The environment is the surroundings of the chemical system, including the container in which the reaction takes place, or the water in which the molecules are dissolved.

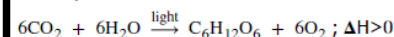
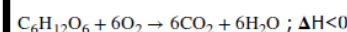
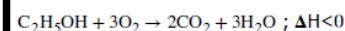
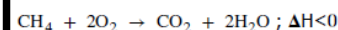
ENDOTHERMIC	EXOTHERMIC
Definition: A reaction that absorbs heat energy from the surroundings	Definition: A reaction that releases heat energy into the environment
More energy absorbed than released	More energy released than absorbed
Net energy change is energy absorbed from the environment	Net energy change is energy released into the environment
The chemical system's energy increases ($\Delta H > 0$)	The chemical system's energy decreases ($\Delta H < 0$)
The environment's energy decreases	The environment's energy increases
Temperature of the environment decreases (test tube gets colder)	Temperature of the environment increases (test tube gets hotter)

**CATALYST**

In order for a reaction to occur, enough energy has to be provided (activation energy) for particles to collide effectively.

The amount of required energy can be decreased by using a catalyst. A catalyst is a chemical substances that lowers the activation energy required without taking part in the reaction. By lowering the activation energy, the rate of the reaction can also be increased.

A catalyst is a substance that increases the rate of the reaction but remains unchanged at the end of the reaction

IMPORTANT REACTIONS**ENDOTHERMIC**Photosynthesis**EXOTHERMIC**Cellular respirationCombustion

IDEAL GASES

KINETIC MOLECULAR THEORY

Properties of gases according to kinetic molecular theory

- Particles are far apart- large spaces between them.
- Particles move with Brownian motion- Fast, random motion.
- Gases undergo diffusion- Particles spread out to fill the container.
- The forces between particles are negligibly small.
- Particles collide elastically with other particles and the walls of the container.

Three measurable variables:

Temperature (T)- a measure of average kinetic energy.

Volume (V)- the space occupied by the gas in the container.

Pressure (p)- the outward force per unit area applied by gas when the gas particles collide with the container walls.

Standard Temperature and Pressure or STP

Standard Temperature = 0 °C or 273 K
Standard Pressure = 101,3 kPa (1,013 x 10⁵ Pa)

IDEAL GASES

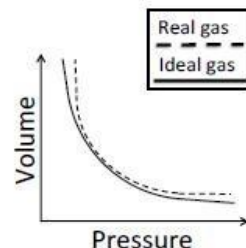
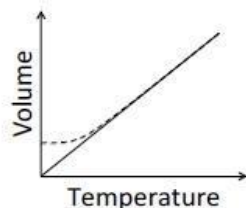
Ideal gases are theoretical gases that obey the gas laws under all conditions of temperature and pressure.

- The particles of a gas are all identical and in constant motion.
- The volume of the gas is due to the motion of the particles as the particles have no volume themselves.
- The intermolecular forces between gas particles are negligible.
- All collisions are perfectly elastic.

DEVIATION FROM IDEAL GAS

At LOW temperatures:

- Gas particles move much slower, decreasing the number of collisions between the particles and the container walls. Pressure of the gas is lower than expected.
- Attractive forces are stronger due to the particles being closer together and moving slower. The gas is more likely to become a liquid and the volume is greater than expected.



At HIGH pressure:

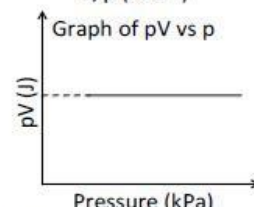
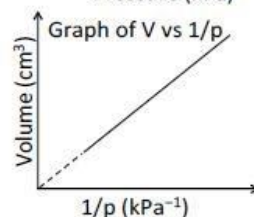
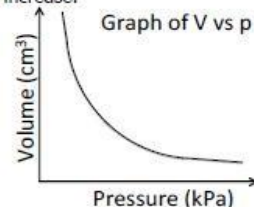
- The volume of the gas particles becomes significant in relation to the container.
- Increases in the intermolecular forces between the particles due to proximity, gas will liquify and the volume is larger than expected.

BOYLE'S LAW

$$p_1 V_1 = p_2 V_2 \quad V \propto \frac{1}{p} \quad (\text{constant } T)$$

The volume of an enclosed gas is inversely proportional to the pressure of the gas if the temperature remains constant.

The temperature remains constant, therefore the average Ek of the particles remains the same. If the volume decreases, more collisions would take place and the pressure would increase.



Example:

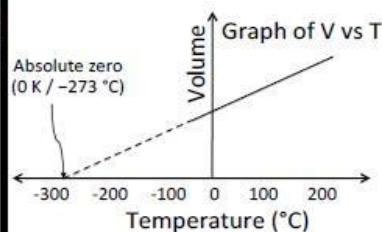
A balloon has a volume of 1000 cm³ at a pressure of 60 kPa. If the pressure on the balloon changes to 100 kPa, calculate the new volume of the balloon.

$$\begin{aligned} p_1 V_1 &= p_2 V_2 \\ (1000)(60) &= (100)V_2 \\ V_2 &= 600 \text{ cm}^3 \end{aligned}$$

CHARLES' LAW

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad V \propto T \quad (\text{constant } P)$$

The volume of an enclosed gas is directly proportional to the temperature (in Kelvin) of the gas provided the pressure remains constant.



By extending the graph, absolute zero (0 Kelvin) can be derived. 0 K is the theoretical minimum temperature, as volume (V) becomes zero at this temperature. Real gases deviate from the ideal gas laws at low temperatures, as the volume of the gas and its particles can never become zero.

$$^{\circ}\text{C} = \text{K} - 273$$

$$\text{K} = ^{\circ}\text{C} + 273$$

NB: Temperature must always be measured in Kelvin!

Example:

A certain sample of gas has a volume of 125 cm³ at standard temperature and pressure. If the pressure of the gas remains constant but the temperature of the gas increases to 25 °C, calculate the new volume of the gas.

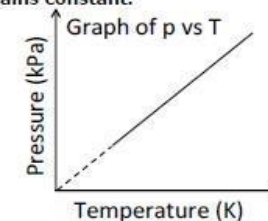
$$T_2 = 25 + 273 = 298 \text{ K}$$

$$\begin{aligned} \frac{V_1}{T_1} &= \frac{V_2}{T_2} \\ \frac{125}{273} &= \frac{V_2}{298} \\ V_2 &= 136,45 \text{ cm}^3 \end{aligned}$$

GAY LUSSAC'S LAW

$$\frac{p_1}{T_1} = \frac{p_2}{T_2} \quad P \propto T \quad (\text{constant } V)$$

The pressure of an enclosed gas is directly proportional to the absolute temperature if the volume of the gas remains constant.



A decrease in temperature indicates a decrease in kinetic energy. As kinetic energy decreases, so does the outward force during collisions between particles and container.

Example:

Calculate the pressure of a gas at 25 °C, if originally the gas had a pressure of 200 Pa at a temperature of 100 °C.

$$\begin{aligned} \frac{p_1}{T_1} &= \frac{p_2}{T_2} \\ \frac{200}{373} &= \frac{p_2}{298} \\ p_2 &= 159,79 \text{ Pa} \end{aligned}$$

GENERAL GAS EQUATION

Variables of fixed amount of gas when system changes occur:

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

p, V can be in any consistent units
T in Kelvin

IDEAL GAS EQUATION

Properties under single set of conditions.

$$pV = nRT$$

p = pressure in Pa

V = volume in m³

n = number of mole of gas (mol)

R = universal gas constant (8,31 J·K⁻¹·mol⁻¹)

T = temperature in K

ACID/BASE DEFINITIONS Arrhenius

An **acid** is a substance that produces hydrogen ions (H^+) / hydronium ions (H_3O^+) when dissolved in water.

A **base** is a substance that produces hydroxide ions (OH^-) when dissolved in water.

Bronsted-Lowry

An **acid** is a proton (H^+ ion) donor.

A **base** is a proton (H^+ ion) acceptor.

ACID PROTICITY

Some acids are able to donate more than one proton. The number of protons that an acid can donate is referred to as the acid proticity.

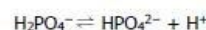
1 proton- monoprotic



2 protons- diprotic



3 protons- triprotic



CONJUGATE ACID-BASE PAIRS

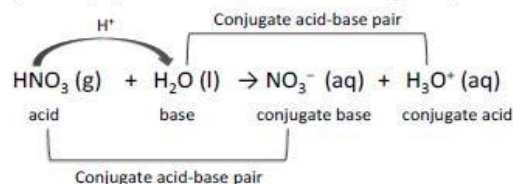
An acid forms a conjugate base when it donates a proton.
acid $\rightleftharpoons$ conjugate base + H^+

A base forms a conjugate acid when it accepts a proton.
base + H^+ $\rightleftharpoons$ conjugate acid

Conjugate acid-base pairs are compounds that differ by the presence of one proton, or H^+ .

EXAMPLE:

Identify the conjugate acid-base pair in the following example:

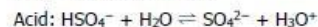


AMPHOLYTE/ AMPHOTERIC SUBSTANCES

Ampholyte- A substance that can act as both an acid or a base.

Amphoteric/amphiprotic substances can therefore either donate or accept protons. Common ampholytes include H_2O , HCO_3^- and HSO_4^- .

HSO_4^- as an ampholyte:



ACIDS AND BASES

CONCENTRATED VS DILUTE ACIDS AND BASES

Concentration is the number of moles of solute per unit volume of solution.

A concentrated acid is an acid with a large amount of solute (the acid) dissolved in a small volume of solvent (water).

A dilute acid is an acid with a small amount of solute (the acid) dissolved in a large volume of solvent (water).

Concentrated strong acid - $1\text{mol}\cdot\text{dm}^{-3}$ of HCl

Concentrated weak acid - $1\text{mol}\cdot\text{dm}^{-3}$ of CH_3COOH

Dilute strong acid - $0,01\text{mol}\cdot\text{dm}^{-3}$ of HCl

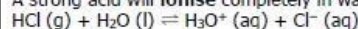
Dilute weak acid - $0,01\text{mol}\cdot\text{dm}^{-3}$ of CH_3COOH

STRONG VS WEAK ACIDS AND BASES

The strength of an acid/base refers to the ability of the substance to ionise or dissociate.

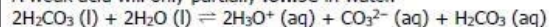
ACIDS

A strong acid will **ionise** completely in water.



(strong acid $\rightarrow$ weak conjugate base)

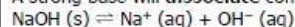
A weak acid will only partially **ionise** in water.



(weak acid $\rightarrow$ strong conjugate base)

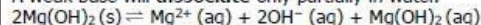
BASES

A strong base will **dissociate** completely in water.



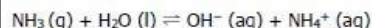
(strong base $\rightarrow$ weak conjugate acid)

A weak base will **dissociate** only partially in water.



(weak base $\rightarrow$ strong conjugate acid)

NH_3 is an exception, it ionises.



INFLUENCE OF ACID/BASE STRENGTH

Reaction rate

Reaction rates increase as the strength of the acid/base increases.

Stronger acid = higher concentration of ions = greater rate of reaction.

Conductivity

Conductivity increases as the strength of the acid/base increases.

Stronger acid = higher concentration of H^+ = greater conductivity.

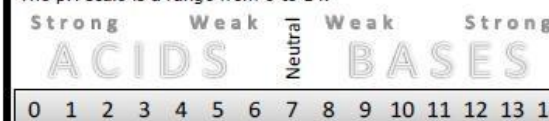
COMMON ACIDS		COMMON BASES	
Hydrochloric acid (HCl)	STRONG	Sodium Hydroxide (NaOH)	STRONG
Nitric acid (HNO ₃)		Potassium hydroxide (KOH)	
Sulfuric acid (H ₂ SO ₄)		Sodium hydrogen carbonate (NaHCO ₃)	
Hydrofluoric acid (HF)	WEAK	Calcium carbonate (CaCO ₃)	WEAK
Sulfurous acid (H ₂ SO ₃)		Sodium carbonate (Na ₂ CO ₃)	
Carbonic acid (H ₂ CO ₃)		Ammonia (NH ₃)	
Acetic acid / ethanoic acid (CH ₃ COOH)			
Phosphoric acid (H ₃ PO ₄)			

THE pH SCALE

The pH of a solution is a number that represents the acidity or alkalinity of a substance.

The greater the concentration of H^+ ions in solution, the more acidic the solution and the lower the pH. The lower the concentration of H^+ in solution, the more alkali the solution and the higher the pH.

The pH scale is a range from 0 to 14.

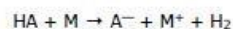


ACIDS AND BASES

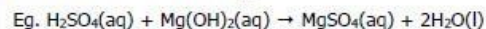
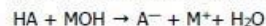
PROTOLYTIC REACTIONS

Protolytic reactions are reactions during which protons (H^+) are transferred.

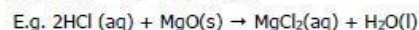
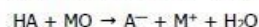
acid + metal $\rightarrow$ salt + hydrogen



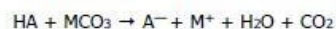
acid + metal hydroxide (base) $\rightarrow$ salt + water



acid + metal oxide $\rightarrow$ salt + water



acid + metal carbonate $\rightarrow$ salt + water + carbon dioxide



The ionic salt is made up of A^- (from the acid) and M^+ (from the base).

HYDROLYSIS OF SALTS

Hydrolysis is the reaction of a salt with water. Salt is a substance in which the hydrogen of an acid has been replaced by a cation. During the hydrolysis, the salt will form an acidic, alkali or neutral solution. The acidity of the salt is dependent on the relative strength of the acid and base that is used.

Strong acid, weak base

The weak base will form a strong conjugate acid. The strong acid will form a weak conjugate base. The strong conjugate acid will form an acidic solution.

Weak acid, strong base

The weak acid will form a strong conjugate base. The strong base will form a weak conjugate acid. The strong conjugate base will form an alkali solution.

ACID STRENGTH	BASE STRENGTH	SALT
Strong	Weak	Acidic
Strong	Strong	Neutral
Weak	Strong	Alkali

INDICATORS

An indicator is a compound that changes colour according to the pH of the substance. During titrations, the indicator needs to be selected according to the acidity/alkalinity of the salt that will be produced (see hydrolysis).

INDICATOR	COLOUR IN ACID	COLOUR IN BASE	COLOUR AT EQUIVALENCE POINT	PH RANGE OF EQUIVALENCE
Litmus	Red	Blue		4,5 - 8,3
Methyl orange	Red	Yellow	Orange	3,1 - 4,4
Bromothymol blue	Yellow	Blue	Green	6,0 - 7,6
Phenolphthalein	Colourless	Pink	Pale Pink	8,3 - 10,0

Neutral: pH 7

Neutralisation is when the equivalence point is reached. Equivalence point is NOT when the solution is at pH 7, but when the molar amount of acid and base is the same according to the molar ratio. The pH at neutralisation is dependent on the salt that is formed. (see hydrolysis).

Neutralisation is the point where an acid and base have reacted so neither is in excess.

TITRATIONS

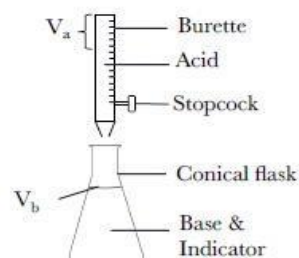
A titration is a practical laboratory method to determine the concentration of an acid or base. The concentration of an acid or base can be determined by accurate neutralisation with a **standard solution**- a solution of known concentration. Neutralisation occurs at the equivalence point, when the molar amount of acid and base is the same according to the molar ratio.

$$\frac{n_a}{n_b} = \frac{c_a V_a}{c_b V_b}$$

n- number of mole of substance (mol)/ mole ratio
c- concentration of acid/base ($\text{mol} \cdot \text{dm}^{-3}$)
V- volume of solution (dm^3)

$$\begin{aligned} 1 \text{ mL} &= 1 \text{ cm}^3 \\ 1 \text{ L} &= 1 \text{ dm}^3 \\ 1000 \text{ mL} &= 1 \text{ L} \\ 1000 \text{ cm}^3 &= 1 \text{ dm}^3 \end{aligned}$$

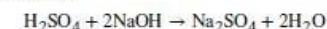
Titration setup



1. Pipette known solution into conical flask (usually base).
2. Add appropriate indicator to flask.
3. Add unknown concentration solution to the burette (usually acid).
4. Add solution from burette to conical flask at a dropwise rate (remember to swirl).
5. Stop burette when indicator shows neutralisation/equivalence point has been reached.

EXAMPLE:

During a titration, 25 cm^3 of dilute H_2SO_4 neutralises 40 cm^3 of NaOH solution. If the concentration of the H_2SO_4 solution is 0,25 $\text{mol} \cdot \text{dm}^{-3}$, calculate the concentration of the NaOH.



$$\begin{aligned} n_a &= 1 & n_b &= 2 \\ c_a &= 0,25 \text{ mol} \cdot \text{dm}^{-3} & c_b &= ? \\ V_a &= 25 \text{ cm}^3 = 0,025 \text{ dm}^3 & V_b &= 40 \text{ cm}^3 = 0,04 \text{ dm}^3 \end{aligned}$$

By titration equations

$$\begin{aligned} \frac{n_a}{n_b} &= \frac{c_a V_a}{c_b V_b} \\ \frac{1}{2} &= \frac{(0,25)(0,025)}{c_b(0,04)} \\ c_b &= 0,31 \text{ mol} \cdot \text{dm}^{-3} \end{aligned}$$

By first principles:

$$\begin{aligned} \text{Acid:} & \\ n &= cV \\ &= (0,25)(0,025) \\ &= 6,25 \times 10^{-3} \text{ mol} \\ \text{Base:} & \\ c &= \frac{n}{V} \\ &= \frac{6,25 \times 10^{-3} \times 2}{0,04} \\ &= 0,31 \text{ mol} \cdot \text{dm}^{-3} \end{aligned}$$