

3. ELABORATION OF THE CONTENT FOR GRADE 12 (CAPS)

The final examination in Physical Sciences will cover the topics outlined below.

3.1 Paper 1: Physics

Newton's laws and application of Newton's laws

(This section must be read in conjunction with the CAPS, p. 62–66.)

Different kinds of forces: weight, normal force, frictional force, applied force (push, pull), tension (strings or cables)

- Define *normal force*, N , as the force or the component of a force which a surface exerts on an object with which it is in contact, and which is perpendicular to the surface.
- Define *frictional force*, f , as the force that opposes the motion of an object and which acts parallel to the surface.
Define *static frictional force*, f_s , as the force that opposes the tendency of motion of a stationary object relative to a surface.
Define *kinetic frictional force*, f_k , as the force that opposes the motion of a moving object relative to a surface.
Know that a frictional force:
 - Is proportional to the normal force
 - Is independent of the area of contact
 - Is independent of the velocity of motion
- Solve problems using $f_s^{\max} = \mu_s N$ where f_s^{\max} is the maximum static frictional force and μ_s is the coefficient of static friction.
NOTE:
 - If a force, F , applied to a body parallel to the surface does not cause the object to move, F is equal in magnitude to the static frictional force.
 - The static frictional force is a maximum (f_s^{\max}) just before the object starts to move across the surface.
 - If the applied force exceeds f_s^{\max} , a resultant/net force accelerates the object.
- Solve problems using $f_k = \mu_k N$, where f_k is the kinetic frictional force and μ_k the coefficient of kinetic friction.

Force diagrams, free-body diagrams

- Draw force diagrams.
- Draw free-body diagrams. (This is a diagram that shows the relative magnitudes and directions of forces acting on a body/particle that has been isolated from its surroundings)
- Resolve a two-dimensional force (such as the weight of an object on an inclined plane) into its parallel (x) and perpendicular (y) components.
- Determine the resultant/net force of two or more forces.

Newton's first, second and third laws of motion

- State Newton's first law of motion: A body will remain in its state of rest or motion at constant velocity unless a non-zero resultant/net force acts on it.
- Discuss why it is important to wear seatbelts using Newton's first law of motion.
- State Newton's second law of motion: When a resultant/net force acts on an object, the object will accelerate in the direction of the force at an acceleration directly proportional to the force and inversely proportional to the mass of the object.
- Draw force diagrams and free-body diagrams for objects that are in equilibrium or accelerating.

- Apply Newton's laws of motion to a variety of equilibrium and non-equilibrium problems including:
 - A single object:
 - Moving on a horizontal plane with or without friction
 - Moving on an inclined plane with or without friction
 - Moving in the vertical plane (lifts, rockets, etc.)
 - Two-body systems (joined by a light inextensible string):
 - Both on a flat horizontal plane with or without friction
 - One on a horizontal plane with or without friction, and a second hanging vertically from a string over a frictionless pulley
 - Both on an inclined plane with or without friction
 - Both hanging vertically from a string over a frictionless pulley
- State Newton's third law of motion: When one body exerts a force on a second body, the second body exerts a force of equal magnitude in the opposite direction on the first body.
- Identify action-reaction pairs.
- List the properties of action-reaction pairs.

Newton's Law of Universal Gravitation

- State Newton's Law of Universal Gravitation: Each body in the universe attracts every other body with a force that is directly proportional to the product of their masses and inversely proportional to the square of the distance between their centres.
- Solve problems using $F = \frac{Gm_1m_2}{r^2}$.
- Calculate acceleration due to gravity on a planet using $g = \frac{Gm}{r^2}$.
- Describe *weight* as the gravitational force the Earth exerts on any object on or near its surface.
- Calculate weight using the expression $w = mg$.
- Calculate the weight of an object on other planets with different values of gravitational acceleration.
- Distinguish between *mass* and *weight*.
- Explain *weightlessness*.

Momentum and Impulse

(This section must be read in conjunction with the CAPS, p. 99–101.)

Momentum

- Define *momentum* as the product of an object's mass and its velocity.
- Describe the *linear momentum* of an object as a vector quantity with the same direction as the velocity of the object.
- Calculate the momentum of a moving object using $p = mv$.
- Describe the *vector nature of momentum* and illustrate it with some simple examples.
- Draw vector diagrams to illustrate the relationship between the initial momentum, the final momentum and the change in momentum for each of the above examples.

Newton's second law of motion in terms of momentum

- State Newton's second law of motion in terms of momentum: The resultant/net force acting on an object is equal to the rate of change of momentum of the object in the direction of the resultant/net force.
- Express Newton's second law of motion in symbols: $F_{\text{net}} = \frac{\Delta p}{\Delta t}$.
- Calculate the change in momentum when a resultant/net force acts on an object and its velocity:
 - Increases in the direction of motion, e.g. 2nd stage rocket engine fires
 - Decreases, e.g. brakes are applied
 - Reverses its direction of motion, e.g. a soccer ball kicked back in the direction it came from

Impulse

- Define *impulse* as the product of the resultant/net force acting on an object and the time the resultant/net force acts on the object.
- Deduce the impulse-momentum theorem: $F_{\text{net}}\Delta t = m\Delta v$.
- Use the impulse-momentum theorem to calculate the force exerted, the time for which the force is applied and the change in momentum for a variety of situations involving the motion of an object in one dimension.
- Explain how the concept of impulse applies to safety considerations in everyday life, e.g. airbags, seatbelts and arrestor beds.

Conservation of momentum and elastic and inelastic collisions

- Explain what is meant by a *closed/an isolated system* (in Physics), i.e. a system on which the resultant/net external force is zero.
A closed/an isolated system excludes external forces that originate outside the colliding bodies, e.g. friction. Only internal forces, e.g. contact forces between the colliding objects, are considered.
- State the principle of conservation of linear momentum: The total linear momentum of a closed system remains constant (is conserved).
- Apply the conservation of momentum to the collision of two objects moving in one dimension (along a straight line) with the aid of an appropriate sign convention.
- Distinguish between *elastic collisions* and *inelastic collisions* by calculation.

Vertical Projectile Motion in One Dimension (1D)

(This section must be read in conjunction with the CAPS, p. 102–103.)

- Explain what is meant by a *projectile*, i.e. an object upon which the only force acting is the force of gravity.
- Use equations of motion to determine the position, velocity and displacement of a projectile at any given time.
- Sketch position versus time (x vs. t), velocity versus time (v vs. t) and acceleration versus time (a vs. t) graphs for:
 - A free-falling object
 - An object thrown vertically upwards
 - An object thrown vertically downwards
 - Bouncing objects (restricted to balls)
- For a given x vs. t , v vs. t or a vs. t graph, determine:
 - Position
 - Displacement
 - Velocity or acceleration at any time t
- For a given x vs. t , v vs. t or a vs. t graph, describe the motion of the object:
 - Bouncing
 - Thrown vertically upwards
 - Thrown vertically downward

Work, Energy and Power

(This section must be read in conjunction with the CAPS, p. 117–120.)

Work

- Define the work done on an object by a constant force F as $F \Delta x \cos \theta$, where F is the magnitude of the force, Δx the magnitude of the displacement and θ the angle between the force and the displacement. (Work is done by a force on an object – the use of 'work is done against a force', e.g. work done against friction, should be avoided.)
- Draw a force diagram and free-body diagrams.
- Calculate the net/total work done on an object.
- Distinguish between *positive net/total work done* and *negative net/total work done* on the system.

Work-energy theorem

- State the work-energy theorem: The net/total work done on an object is equal to the change in the object's kinetic energy OR the work done on an object by a resultant/net force is equal to the change in the object's kinetic energy.
In symbols: $W_{\text{net}} = \Delta K = K_f - K_i$.
- Apply the work-energy theorem to objects on horizontal, vertical and inclined planes (for both frictionless and rough surfaces).

Conservation of energy with non-conservative forces present

- Define a *conservative force* as a force for which the work done in moving an object between two points is independent of the path taken. Examples are gravitational force, the elastic force in a spring and electrostatic forces (coulomb forces).
- Define a *non-conservative force* as a force for which the work done in moving an object between two points depends on the path taken. Examples are frictional force, air resistance, tension in a chord, etc.
- State the principle of conservation of mechanical energy: The total mechanical energy (sum of gravitational potential energy and kinetic energy) in an isolated system remains constant. (A system is isolated when the resultant/net external force acting on the system is zero.)
- Solve conservation of energy problems using the equation: $W_{\text{nc}} = \Delta E_k + \Delta E_p$
- Use the relationship above to show that in the absence of non-conservative forces, mechanical energy is conserved.

Power

- Define *power* as the rate at which work is done or energy is expended.
In symbols: $P = \frac{W}{\Delta t}$
- Calculate the power involved when work is done.
- Perform calculations using $P_{ave} = Fv_{ave}$ when an object moves at a constant speed along a rough horizontal surface or a rough inclined plane.
- Calculate the power output for a pump lifting a mass (e.g. lifting water through a height at constant speed).

Doppler Effect (relative motion between source and observer)

(This section must be read in conjunction with the CAPS, p. 121–122.)

With sound and ultrasound

- State the Doppler effect as the change in frequency (or pitch) of the sound detected by a listener because the sound source and the listener have different velocities relative to the medium of sound propagation.
- Explain (using appropriate illustrations) the change in pitch observed when a source moves toward or away from a listener.
- Solve problems using the equation $f_L = \frac{v \pm v_L}{v \pm v_s} f_s$ when EITHER the source or the listener is moving.
- State applications of the Doppler effect.

With light – red shifts in the universe (evidence for the expanding universe)

- Explain *red shifts* and *blue shifts* using the Doppler Effect.
- Use the Doppler effect to explain why we conclude that the universe is expanding.

Electrostatics

(This section must be read in conjunction with the CAPS, p. 84–85.)

Coulomb's law

- State Coulomb's law: The magnitude of the electrostatic force exerted by one point charge (Q_1) on another point charge (Q_2) is directly proportional to the product of the magnitudes of the charges and inversely proportional to the square of the distance (r) between them:
- Solve problems using the equation $F = \frac{kQ_1Q_2}{r^2}$ for charges in one dimension (1D) (restrict to three charges).
- Solve problems using the equation $F = \frac{kQ_1Q_2}{r^2}$ for charges in two dimensions (2D) – for three charges in a right-angled formation (limit to charges at the 'vertices of a right-angled triangle').

Electric field

- Describe an *electric field* as a region of space in which an electric charge experiences a force. The direction of the electric field at a point is the direction that a positive test charge would move if placed at that point.
- Draw electric field patterns for the following configurations:
 - A single point charge
 - Two point charges (one negative, one positive OR both positive OR both negative)
 - A charged sphere

NOTE: Restrict to situations in which the charges are identical in magnitude.

- Define the *electric field at a point*. The electric field at a point is the electrostatic force experienced per unit positive charge placed at that point. In symbols: $E = \frac{F}{q}$.
- Solve problems using the equation $E = \frac{F}{q}$.
- Calculate the electric field at a point due to a number of point charges, using the equation $E = \frac{kQ}{r^2}$ to determine the contribution to the field due to each charge. Restrict to three charges in a straight line.

Electric Circuits

(This section must be read in conjunction with the CAPS, p. 88–89 & 121.)

Ohm's law

- State Ohm's law in words: The potential difference across a conductor is directly proportional to the current in the conductor at constant temperature.
- Determine the relationship between current, potential difference and resistance at constant temperature using a simple circuit.
- State the difference between *ohmic conductors* and *non-ohmic conductors* and give an example of each.
- Solve problems using $R = \frac{V}{I}$ for series and parallel circuits (maximum four resistors).

Power, energy

- Define *power* as the rate at which work is done.
- Solve problems using $P = \frac{W}{\Delta t}$.
- Solve problems using $P = VI$, $P = I^2R$ or $P = \frac{V^2}{R}$.
- Solve circuit problems involving the concepts of power and electrical energy.
- Deduce that the kilowatt hour (kWh) refers to the use of 1 kilowatt of electricity for 1 hour.
- Calculate the cost of electricity usage given the power specifications of the appliances used, the duration and the cost of 1 kWh.

Internal resistance, series and parallel networks

- Solve problems involving current, voltage and resistance for circuits containing arrangements of resistors in series and in parallel (maximum four resistors).
- Explain the term *internal resistance*.
- Solve circuit problems using $\varepsilon = V_{\text{load}} + V_{\text{internal resistance}}$ or $\varepsilon = IR_{\text{ext}} + Ir$.
- Solve circuit problems, with internal resistance, involving series-parallel networks of resistors (maximum four resistors).

Electrodynamics

(This section must be read in conjunction with the CAPS, p. 130–131.)

Electrical machines (generators, motors)

- State the energy conversion in generators.
- Use the principle of electromagnetic induction to explain how a generator works.
- Explain the functions of the components of an AC and a DC generator.
- State examples of the uses of AC and DC generators.
- State the energy conversion in motors.

- Use the motor effect to explain how a motor works.
- Explain the functions of the components of a motor.
- State examples of the use of motors.

Alternating current

- State the advantages of alternating current over direct current.
- Sketch graphs of voltage versus time and current versus time for an AC circuit.
- Define the term *rms* for an alternating voltage or an alternating current. The rms value of AC is the DC potential difference/current which dissipates the same amount of energy as AC.

- Solve problems using $I_{\text{rms}} = \frac{I_{\text{max}}}{\sqrt{2}}$, $V_{\text{rms}} = \frac{V_{\text{max}}}{\sqrt{2}}$.

- Solve problems using $P_{\text{ave}} = I_{\text{rms}} V_{\text{rms}} = \frac{1}{2} I_{\text{max}} V_{\text{max}}$ (for a purely resistive circuit),

$$P_{\text{ave}} = I_{\text{rms}}^2 R \text{ and } P_{\text{ave}} = \frac{V_{\text{rms}}^2}{R}$$

Optical Phenomena and Properties of Materials

(This section must be read in conjunction with the CAPS, p. 132–133.)

Photo-electric effect

- Describe the *photoelectric effect* as the process whereby electrons are ejected from a metal surface when light of suitable frequency is incident on that surface.
- State the significance of the photoelectric effect.
- Define *threshold frequency*, f_0 , as the minimum frequency of light needed to emit electrons from a certain metal surface.
- Define *work function*, W_0 , as the minimum energy that an electron in the metal needs to be emitted from the metal surface.
- Perform calculations using the photoelectric equation:
 $E = W_0 + K_{\text{max}}$, where $E = hf$ and $W_0 = hf_0$ and $K_{\text{max}} = \frac{1}{2}mv_{\text{max}}^2$
- Explain the effect of intensity and frequency on the photoelectric effect.

Emission and absorption spectra

- Explain the *formation of atomic spectra* by referring to energy transition.
- Explain the difference between *atomic absorption spectra* and *atomic emission spectra*.
An atomic absorption spectrum is formed when certain frequencies of electromagnetic radiation that passes through a medium, e.g. a cold gas, is absorbed.
An atomic emission spectrum is formed when certain frequencies of electromagnetic radiation are emitted due to an atom's electrons making a transition from a high-energy state to a lower energy state.

3.2 Paper 2: Chemistry

Representing Chemical Change

(This section must be read in conjunction with the CAPS, p. 37.)

Balanced chemical equations

- Write and balance chemical equations.
- Interpret balanced reaction equations in terms of:
 - Conservation of atoms
 - Conservation of mass (use relative atomic masses)

Quantitative Aspects of Chemical Change

(This section must be read in conjunction with the CAPS, p. 82.)

Molar volume of gases

- 1 mole of any gas occupies 22,4 dm³ at 0 °C (273 K) and 1 atmosphere (101,3 kPa).

Volume relationships in gaseous reactions

- Interpret balanced equations in terms of volume relationships for gases, i.e. under the same conditions of temperature and pressure, equal number of moles of all gases occupy the same volume.

Concentration of solutions

- Calculate the molar concentration of a solution.

More complex stoichiometric calculations

- Determine the empirical formula and molecular formula of compounds.
- Determine the percentage yield of a chemical reaction.
- Determine percentage purity or percentage composition, e.g. the percentage CaCO₃ in an impure sample of seashells.
- Perform stoichiometric calculations based on balanced equations.
- Perform stoichiometric calculations based on balanced equations that may include limiting reagents.

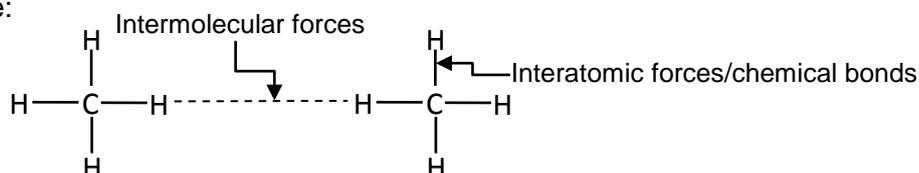
Intermolecular Forces

(This section must be read in conjunction with the CAPS, p. 71–73.)

Intermolecular forces and interatomic forces (chemical bonds)

- Name and explain the different intermolecular forces (Van der Waal's forces):
 - (i) Dipole-dipole forces:
Forces between two polar molecules
 - (ii) Induced dipole forces or London forces:
Forces between non-polar molecules
 - (iii) Hydrogen bonding:
Forces between molecules in which hydrogen is covalently bonded to nitrogen, oxygen or fluorine – a special case of dipole-dipole forces
- Describe the difference between *intermolecular forces* and *interatomic forces* (intramolecular forces) using a diagram of a group of small molecules; and in words.

Example:



- State the relationship between intermolecular forces and molecular size. For non-polar molecules, the strength of induced dipole forces increases with molecular size.

- Explain the effect of intermolecular forces on boiling point, melting point and vapour pressure.
Boiling point:
The temperature at which the vapour pressure of a substance equals atmospheric pressure. The stronger the intermolecular forces, the higher the boiling point.
Melting point:
The temperature at which the solid and liquid phases of a substance are at equilibrium. The stronger the intermolecular forces, the higher the melting point.
Vapour pressure:
The pressure exerted by a vapour at equilibrium with its liquid in a closed system. The stronger the intermolecular forces, the lower the vapour pressure.

Organic Molecules

(This section must be read in conjunction with the CAPS, p. 104–116.)

- Define *organic molecules* as molecules containing carbon atoms.

Organic molecular structures – functional groups, saturated and unsaturated structures, isomers

- Write down condensed structural formulae, structural formulae and molecular formulae (up to 8 carbon atoms, one functional group per molecule) for:
 - Alkanes (no ring structures)
 - Alkenes (no ring structures)
 - Alkynes
 - Halo-alkanes (primary, secondary and tertiary haloalkanes; no ring structures)
 - Alcohols (primary, secondary and tertiary alcohols)
 - Carboxylic acids
 - Esters
 - Aldehydes
 - Ketones

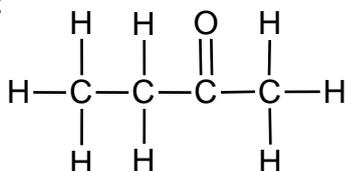
- Know the following definitions/terms:

Molecular formula: A chemical formula that indicates the type of atoms and the correct number of each in a molecule.

Example: C₄H₈O

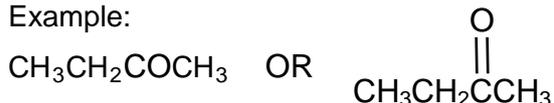
Structural formula: A structural formula of a compound shows which atoms are attached to which within the molecule. Atoms are represented by their chemical symbols and lines are used to represent ALL the bonds that hold the atoms together.

Example:



Condensed structural formula: This notation shows the way in which atoms are bonded together in the molecule, but DOES NOT SHOW ALL bond lines.

Example:



Hydrocarbon: Organic compounds that consist of hydrogen and carbon only.

Homologous series: A series of organic compounds that can be described by the same general formula OR in which one member differs from the next with a CH₂ group.

Saturated compounds: Compounds in which there are no multiple bonds between C atoms in their hydrocarbon chains.

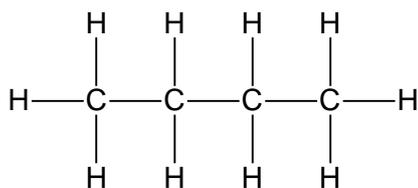
Unsaturated compounds: Compounds with one or more multiple bonds between C atoms in their hydrocarbon chains.

Functional group: A bond or an atom or a group of atoms that determine(s) the physical and chemical properties of a group of organic compounds.

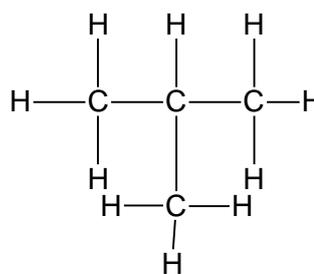
Homologous series	Structure of functional group	
	Structure	Name/Description
Alkanes	$\begin{array}{c} & \\ -C & -C- \\ & \end{array}$	Only C-H and C-C single bonds
Alkenes	$\begin{array}{c} \diagup & \diagdown \\ C & =C \\ \diagdown & \diagup \end{array}$	Carbon-carbon double bond
Alkynes	$-C \equiv C-$	Carbon-carbon triple bond
Haloalkanes	$\begin{array}{c} \\ -C-X \\ \\ (X = F, Cl, Br, I) \end{array}$	Halogen atom bonded to a saturated C atom
Alcohols	$\begin{array}{c} \\ -C-O-H \\ \end{array}$	Hydroxyl group bonded to a saturated C atom
Aldehydes	$\begin{array}{c} O \\ \\ -C-H \end{array}$	Formyl group
Ketones	$\begin{array}{c} & O & \\ & & \\ -C & -C & -C- \\ & & \end{array}$	Carbonyl group bonded to two C atoms
Carboxylic acids	$\begin{array}{c} O \\ \\ -C-O-H \end{array}$	Carboxyl group
Esters	$\begin{array}{c} O \\ \\ -C-O-C- \\ \end{array}$	-

Structural isomer: Organic molecules with the same molecular formula, but different structural formulae.

- Identify compounds (up to 8 carbon atoms) that are saturated, unsaturated and are structural isomers.
- Restrict structural isomers to chain isomers, positional isomers and functional isomers.
 - Chain isomers: Same molecular formula, but different types of chains, e.g. butane and 2-methylpropane.

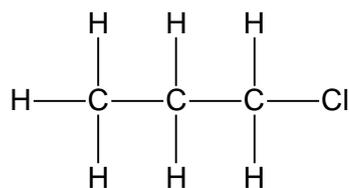


Butane

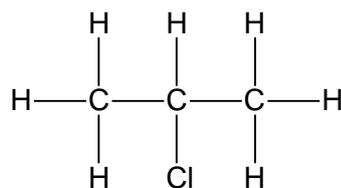


2-methylpropane

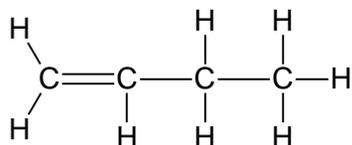
- Positional isomers: Same molecular formula, but different positions of the side chain, substituents or functional groups on the parent chain, e.g. 1-chloropropane and 2-chloropropane or but-2-ene and but-1-ene



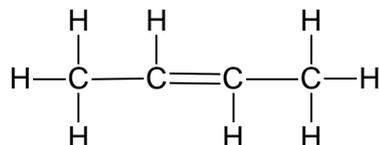
1-chloropropane



2-chloropropane

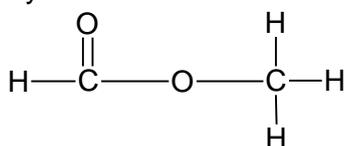


but-1-ene

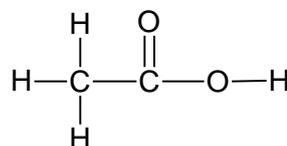


but-2-ene

- Functional isomers: Same molecular formula, but different functional groups, e.g. methyl methanoate and ethanoic acid.



methyl methanoate



ethanoic acid

IUPAC naming and formulae

- Write down the IUPAC name when given the structural formula or condensed structural formula for compounds from the homologous series above, restricted to one functional group per compound, except for haloalkanes. For haloalkanes, maximum two functional groups per molecule.
- Write down the structural formula when given the IUPAC name for the above homologous series.
- Identify alkyl substituents (methyl- and ethyl-) in a chain to a maximum of THREE alkyl substituents on the parent chain.
- When naming haloalkanes, the halogen atoms do not get preference over alkyl groups – numbering should start from the end nearest to the first substituent, either the alkyl group or the halogen. In haloalkanes, where e.g. a Br and a Cl have the same number when numbered from different ends of chain, Br gets alphabetical preference.
- When writing IUPAC names, substituents appear as prefixes written alphabetically (bromo, chloro, ethyl, methyl), ignoring the prefixes di- and tri.

Structure and physical properties (boiling point, melting point, vapour pressure) relationships

- For a given example (from the above functional groups), explain the relationship between physical properties and:
 - Strength of intermolecular forces (Van der Waal's forces), i.e. hydrogen bonds, dipole-dipole forces, induced dipole forces
 - Type of functional groups
 - Chain length
 - Branched chains

Oxidation of alkanes

- State the use of alkanes as fuels.
- Write down an equation for the combustion of an alkane in excess oxygen.

Esterification

- Write down an equation, using structural formulae, for the formation of an ester.
- Name the alcohol and carboxylic acid used and the ester formed.
- Write down reaction conditions for esterification.

Substitution, addition and elimination reactions

- Identify reactions as elimination, substitution or addition.
- Write down, using structural formulae, equations and reaction conditions for the following addition reactions of alkenes:
 - Hydrohalogenation:
The addition of a hydrogen halide to an alkene
 - Halogenation:
The reaction of a halogen (Br_2 , Cl_2) with a compound
 - Hydration:
The addition of water to a compound
 - Hydrogenation:
The addition of hydrogen to an alkene
- Write down, using structural formulae, equations and reaction conditions for the following elimination reactions:
 - Dehydrohalogenation of haloalkanes:
The elimination of hydrogen and a halogen from a haloalkane
 - Dehydration of alcohols:
Elimination of water from an alcohol
 - Cracking of alkanes:
The chemical process in which longer chain hydrocarbon molecules are broken down to shorter more useful molecules.
- Write down, using structural formulae, equations and reaction conditions for the following substitution reactions:
 - Hydrolysis of haloalkanes
Hydrolysis: The reaction of a compound with water
 - Reactions of HX ($\text{X} = \text{Cl}, \text{Br}$) with alcohols to produce haloalkanes
 - Halogenation of alkanes
The reaction of a halogen (Br_2 , Cl_2) with a compound
- Distinguish between *saturated* and *unsaturated hydrocarbons* using bromine water.

Plastics and polymers (ONLY BASIC POLYMERISATION as application of organic chemistry)

- Describe the following terms:
Macromolecule: A molecule that consists of a large number of atoms
Polymer: A large molecule composed of smaller monomer units covalently bonded to each other in a repeating pattern
Monomer: Small organic molecules that can be covalently bonded to each other in a repeating pattern
Polymerisation: A chemical reaction in which monomer molecules join to form a polymer
- Distinguish between *addition polymerisation* and *condensation polymerisation*:
Addition polymerisation: A reaction in which small molecules join to form very large molecules by adding on double bonds
Addition polymer: A polymer formed when monomers (usually containing a double bond) combine through an addition reaction
Condensation polymerisation: Molecules of two monomers with different functional groups undergo condensation reactions with the loss of small molecules, usually water

Condensation polymer: A polymer formed by two monomers with different functional groups that are linked together in a condensation reaction in which a small molecule, usually water, is lost

- Identify monomers from given addition polymers.
- Write down an equation for the polymerisation of ethene to produce polythene.
- State the industrial uses of polythene.

Energy and Change

(This section must be read in conjunction with the CAPS, p. 90–91.)

Energy changes in reactions related to bond energy changes

- Define *heat of reaction* (ΔH) as the energy absorbed or released in a chemical reaction.
- Define *exothermic reactions* as reactions that release energy.
- Define *endothermic reactions* as reactions that absorb energy.
- Classify (with reason) reactions as exothermic or endothermic.

Exothermic and endothermic reactions

- State that $\Delta H > 0$ for endothermic reactions, i.e. reactions in which energy is released.
- State that $\Delta H < 0$ for exothermic reactions, i.e. reactions in which energy is absorbed.

Activation energy

- Define *activation energy* as the minimum energy needed for a reaction to take place.
- Define an *activated complex* as the unstable transition state from reactants to products.
- Draw or interpret fully labelled sketch graphs (potential energy versus course of reaction graphs) of catalysed and uncatalysed endothermic and exothermic reactions.

Rate and Extent of Reaction

(This section must be read in conjunction with the CAPS, p. 123–124.)

Rates of reaction and factors affecting rate

- Define *reaction rate* as the change in concentration of reactants or products per unit time.
- Calculate reaction rate from given data.

$$\text{Rate} = \frac{\Delta C}{\Delta t} \quad (\text{Unit: mol}\cdot\text{dm}^{-3}\cdot\text{s}^{-1})$$

Questions may also include calculations of rate in terms of change in mass/volume/number of moles per time.

- List the factors that affect the rate of chemical reactions, i.e. nature of reacting substances, surface area, concentration (pressure for gases), temperature and the presence of a catalyst.
- Explain in terms of the collision theory how the various factors affect the rate of chemical reactions. The collision theory is a model that explains reaction rate as the result of particles colliding with a certain minimum energy.

Measuring rates of reaction

- Answer questions and interpret data (tables or graphs) on different experimental techniques for measuring the rate of a given reaction.

Mechanism of reaction and of catalysis

- Define the term *positive catalyst* as a substance that increases the rate of a chemical reaction without itself undergoing a permanent change.

- Interpret graphs of distribution of molecular energies (number of particles against their kinetic energy or Maxwell-Boltzmann curves) to explain how a catalyst, temperature and concentration affect rate.
- Explain that a catalyst increases the rate of a reaction by providing an alternative path of lower activation energy. It therefore decreases the net/total activation energy.

Chemical Equilibrium

(This section must be read in conjunction with the CAPS, p. 125–126.)

Chemical equilibrium and factors affecting equilibrium

- Explain what is meant by:
 - Open and closed systems: An open system continuously interacts with its environment, while a closed system is isolated from its surroundings.
 - A reversible reaction: A reaction is reversible when products can be converted back to reactants.
 - Chemical equilibrium: It is a dynamic equilibrium when the rate of the forward reaction equals the rate of the reverse reaction.
- List the factors that influence the position of an equilibrium, i.e. pressure (gases only), concentration and temperature.

Equilibrium constant

- List the factors that influence the value of the equilibrium constant, K_c .
- Write down an expression for the equilibrium constant having been given the equation for the reaction.
- Perform calculations based on K_c values.
- Explain the significance of high and low values of the equilibrium constant.

Application of equilibrium principles

- State Le Chatelier's principle: When the equilibrium in a closed system is disturbed, the system will re-instate a new equilibrium by favouring the reaction that will oppose the disturbance.
- Use Le Chatelier's principle to explain changes in equilibria qualitatively.
- Interpret graphs of equilibrium, e.g. concentration/rate/number of moles/mass/volume versus time graphs.
- Explain the use of rate and equilibrium principles in the Haber process and the contact process.

Acids and Bases

(This section must be read in conjunction with the CAPS, p. 127–128.)

Acid-base reactions

- Define *acids* and *bases* according to Arrhenius and Lowry-Brønsted:
Arrhenius theory: An acid is a substance that produces hydrogen ions (H^+)/hydronium ions (H_3O^+) when it dissolves in water. A base is a substance that produces hydroxide ions (OH^-) when it dissolves in water.
Lowry-Brønsted theory: An acid is a proton (H^+ ion) donor. A base is a proton (H^+ ion) acceptor.
- Distinguish between *strong acids/bases* and *weak acids/bases* with examples.
Strong acids ionise completely in water to form a high concentration of H_3O^+ ions. Examples of strong acids are hydrochloric acid, sulphuric acid and nitric acid.
Weak acids ionise incompletely in water to form a low concentration of H_3O^+ ions. Examples of weak acids are ethanoic acid and oxalic acid.
Strong bases dissociate completely in water to form a high concentration of OH^- ions. Examples of strong bases are sodium hydroxide and potassium hydroxide.
Weak bases dissociate/ionise incompletely in water to form a low concentration of OH^- ions.

Examples of weak bases are ammonia, calcium carbonate, potassium carbonate, calcium carbonate and sodium hydrogen carbonate.

- Distinguish between *concentrated acids/bases* and *dilute acids/bases*.
Concentrated acids/bases contain a large amount (number of moles) of acid/base in proportion to the volume of water.
Dilute acids/bases contain a small amount (number of moles) of acid/base in proportion to the volume of water.
 - Write down the reaction equations of aqueous solutions of acids and bases.
Examples: $\text{HCl}(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ (HCl is a monoprotic acid.)
 $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
 $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell) \rightarrow 2\text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$ (H_2SO_4 is a diprotic acid.)
 - Identify conjugate acid-base pairs for given compounds. When the acid, HA, loses a proton, its conjugate base, A^- , is formed. When the base, A^- , accepts a proton, its conjugate acid, HA, is formed. These two are a conjugate acid-base pair.
 - Describe a substance that can act as either acid or base as amphiprotic or as an ampholyte. Water is a good example of an ampholyte. Write equations to show how an amphiprotic substance can act as acid or base.
 - Write down neutralisation reactions of common laboratory acids and bases.
Examples: $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq})/\text{KOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq})/\text{KCl}(\text{aq}) + \text{H}_2\text{O}(\ell)$
 $\text{HCl}(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g})$
 $\text{HNO}_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{H}_2\text{O}(\ell)$
 $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell)$
 $(\text{COOH})_2(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow (\text{COO})_2\text{Na}_2(\text{aq}) + \text{H}_2\text{O}(\ell)$
 $\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{CH}_3\text{COONa}(\text{aq}) + \text{H}_2\text{O}(\ell)$
- NOTE:** The above are examples of equations that candidates should be able to write from given information. However, any other neutralisation reaction can be given in the question paper to assess, e.g., stoichiometry calculations.
- Determine the approximate pH (equal to, smaller than or larger than 7) of salts in salt hydrolysis. Define *hydrolysis* as the reaction of a salt with water.
 - Hydrolysis of the salt of a weak acid and a strong base results in an alkaline solution, i.e. the $\text{pH} > 7$. Examples of such salts are sodium ethanoate, sodium oxalate and sodium carbonate.
 - Hydrolysis of the salt of a strong acid and a weak base results in an acidic solution, i.e. the $\text{pH} < 7$. An example of such a salt is ammonium chloride.
 - The salt of a strong acid and a strong bases does not undergo hydrolysis and the solution of the salt will be neutral, i.e. $\text{pH} = 7$.
 - Motivate the choice of a specific indicator in a titration. Choose from methyl orange, phenolphthalein and bromothymol blue. Define the *equivalence point* of a titration as the point at which the acid /base has completely reacted with the base/acid.
Define the *endpoint* of a titration as the point where the indicator changes colour.
 - Perform stoichiometric calculations based on titrations of a strong acid with a strong base, a strong acid with a weak base and a weak acid with a strong base. Calculations may include percentage purity.
 - For a titration, e.g. the titration of oxalic acid with sodium hydroxide:
 - List the apparatus needed or identify the apparatus from a diagram.
 - Describe the procedure to prepare a standard oxalic acid solution.
 - Describe the procedure to conduct the titration.
 - Describe safety precautions.
 - Describe measures that need to be in place to ensure reliable results.
 - Interpret given results to determine the unknown concentration.
 - Explain the pH scale as a scale of numbers from 0 to 14 used to express the acidity or alkalinity of a solution.
 - Calculate pH values of strong acids and strong bases using $\text{pH} = -\log[\text{H}_3\text{O}^+]$.

- Define K_w as the equilibrium constant for the ionisation of water or the ionic product of water or the ionisation constant of water, i.e. $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$ by 298 K.
- Explain the *auto-ionisation of water*, i.e. the reaction of water with itself to form H_3O^+ ions and OH^- ions.
- Interpret K_a values of acids to determine the relative strength of given acids. Interpret K_b values of bases to determine the relative strength of given bases.
- Compare strong and weak acids by looking at:
 - pH (monoprotic and diprotic acids)
 - Conductivity
 - Reaction rate

Electrochemical Reactions

(This section must be read in conjunction with the CAPS, p. 134–137.)

Electrolytic cells and galvanic cells

- Define the *galvanic cell* as a cell in which chemical energy is converted into electrical energy. A galvanic (voltaic) cell has self-sustaining electrode reactions.
- Define the *electrolytic cell* as a cell in which electrical energy is converted into chemical energy.
- Define oxidation and reduction in terms of electron (e^-) transfer:
Oxidation is a loss of electrons. Reduction is a gain of electrons.
- Define oxidation and reduction in terms of oxidation numbers:
Oxidation: An increase in oxidation number
Reduction: A decrease in oxidation number
- Define an *oxidising agent* and a *reducing agent* in terms of oxidation and reduction:
Oxidising agent: A substance that is reduced/gains electrons.
Reducing agent: A substance that is oxidised/loses electrons.
- Define an *anode* and a *cathode* in terms of oxidation and reduction:
Anode: The electrode where oxidation takes place
Cathode: The electrode where reduction takes place
- Define an *electrolyte* as a solution/liquid/dissolved substance that conducts electricity through the movement of ions
- Electrolysis: The chemical process in which electrical energy is converted to chemical energy OR the use of electrical energy to produce a chemical change.

Relation of current and potential difference to rate and equilibrium

- Give and explain the relationship between current in an electrolytic cell and the rate of the reaction.
- State that the potential difference of a galvanic cell (V_{cell}) is related to the extent to which the spontaneous cell reaction has reached equilibrium.
- State and use the qualitative relationship between V_{cell} and the concentration of product ions and reactant ions for the spontaneous reaction, namely V_{cell} decreases as the concentration of product ions increases and the concentration of reactant ions decreases until equilibrium is reached at which the $V_{\text{cell}} = 0$ (the cell is 'flat'). (Qualitative treatment only. Nernst equation is NOT required.)

Understanding of the processes and redox reactions taking place in galvanic cells

- Describe the movement of ions in the solutions.
- State the direction of electron flow in the external circuit.
- Write down the half-reactions that occur at the electrodes.
- State the function of the salt bridge.
- Use cell notation or diagrams to represent a galvanic cell.

When writing cell notation, the following convention should be used:

- The $\text{H}_2|\text{H}^+$ half-cell is treated just like any other half-cell.
- Cell terminals (electrodes) are written on the outside of the cell notation.
- Active electrodes:
reducing agent | oxidised species || oxidising agent | reduced species
- Inert electrodes (usually Pt or C):
Pt | reducing agent | oxidised species || oxidising agent | reduced species | Pt
Example: $\text{Pt} | \text{Cl}^-(\text{aq}) | \text{Cl}_2(\text{g}) || \text{F}_2(\text{g}) | \text{F}^-(\text{aq}) | \text{Pt}$
- Predict the half-cell in which oxidation will take place when two half-cells are connected.
- Predict the half-cell in which reduction will take place when connected to another half-cell.
- Write down the overall cell reaction by combining two half-reactions.
- Use the Table of Standard Reduction Potentials to calculate the emf of a standard galvanic cell.
- Use a positive value of the standard emf as an indication that the reaction is spontaneous under standard conditions.

Standard electrode potentials

- Write down the standard conditions under which standard electrode potentials are determined.
- Describe the standard hydrogen electrode and explain its role as the reference electrode.
- Explain how standard electrode potentials can be determined using the reference electrode and state the convention regarding positive and negative values.

Understanding the processes and redox reactions taking place in electrolytic cells

- Describe the movement of ions in the solution.
- State the direction of electron flow in the external circuit.
- Write equations for the half-reactions taking place at the anode and cathode.
- Write down the overall cell reaction by combining two half-reactions.
- Describe, using half-reactions and the equation for the overall cell reaction as well as the layout of the particular cell using a schematic diagram, the following electrolytic processes:
 - The decomposition of copper(II) chloride
 - Electroplating, e.g. the electroplating of an iron spoon with silver/nickel
 - Refining of copper
 - The electrolysis of a concentrated solution of sodium chloride and its use in the chlor-alkali industry
 - The recovery of aluminium metal from bauxite (South Africa uses bauxite from Australia.)
- Describe risks to the environment of the following electrolytic processes used industrially:
 - The production of chlorine (the chemical reactions of the chloro-alkali industry)
 - The recovery of aluminium metal from bauxite

Chemical Industry

(This section must be read in conjunction with the CAPS, p. 138–140.)

The fertiliser industry (N, P, K)

- List, for plants:
 - Three non-mineral nutrients C, H and O and their sources, i.e. the atmosphere (CO_2) and rain (H_2O)
 - Three primary nutrients N, P and K and their sources

- Explain why fertilisers are needed.
- Explain the function of N, P and K in plants.
- Interpret the N : P : K fertiliser ratio and perform calculations based on the ratio.
- Describe, explain, write balanced equations and interpret flow diagrams of the following processes in the industrial manufacture of fertilisers:
 - N_2 – fractional distillation of air
 - H_2 – at SASOL from coal and steam
 - NH_3 – Haber process
 - HNO_3 – Ostwald process
 - H_2SO_4 – Contact process
 - NH_4NO_3 ; $(NH_4)_2SO_4$
- Evaluate the use of inorganic fertilisers on humans and the environment.
- Define *eutrophication* as the process by which an ecosystem, e.g. a river or dam, becomes enriched with inorganic plant nutrients, especially phosphorus and nitrogen, resulting in excessive plant growth. As plant growth becomes excessive, the amount of dead and decaying plant material increases rapidly.
- Discuss alternatives to inorganic fertilisers as used by some communities.