PHYSICAL SCIENCES

EXAMINATION GUIDELINES

GRADE 10

2015

These guidelines consist of 34 pages.

TABLE OF CONTENTS

1.   INTRODUCTION

The Curriculum and Assessment Policy Statement (CAPS) for Physical Sciences outlines the nature and purpose of the subject Physical Sciences. This guides the philosophy underlying the teaching and assessment of the subject in Grade 10.

The purpose of these Examination Guidelines is to:

- Provide clarity on the depth and scope of the content to be assessed in the Grade 10 common/national examination in Physical Sciences.
- Assist teachers to adequately prepare learners for the examinations.

This document deals with the final Grade 10 examinations. It does not deal in any depth with the School-Based Assessment (SBA).

These Examination Guidelines should be read in conjunction with:

- *The National Curriculum Statement (NCS) Curriculum and Assessment Policy Statement (CAPS): Physical Sciences*
- The National Protocol of Assessment: *An addendum to the policy document, the National Senior Certificate: A qualification at Level 4 on the National Qualifications Framework (NQF), regarding the National Protocol for Assessment (Grades R–12)*
- The national policy pertaining to the programme and promotion requirements of the National Curriculum Statement, Grades R–12

2. ASSESSMENT IN GRADE 10

2.1 Format of question papers

Paper	Types of questions	Duration	Total	Date	Marking
1	**Physics** 10 multiple-choice questions – 20 marks Structured questions – 130 marks	2 hours	150	November	Internal
2	**Chemistry** 10 multiple-choice questions – 20 marks Structured questions – 130 marks	2 hours	150	November	Internal

2.2 Numbering and sequence of questions

QUESTION 1: Multiple-choice questions
Subquestions numbered 1.1 to 1.10 (2 marks each)
Questions will be set across all cognitive levels and arranged from lower to higher cognitive levels.

QUESTION 2 onwards:
Longer questions that will assess skills and knowledge across cognitive levels. Numbering starts with QUESTION 2 and will be continuous. Subquestions will be numbered by two digits, e.g. 2.1, 2.2. Numbering is restricted to a maximum of three digits, e.g. 2.1.1, 2.1.2.

2.3 Information sheets

The separate information sheets for Paper 1 and Paper 2 are included in this document.

2.4 Weighting of cognitive levels

Papers 1 and 2 will include questions across all four cognitive levels. The distribution of cognitive levels in Physics and Chemistry papers is given below.

Cognitive level	Description	Paper 1 (Physics)	Paper 2 (Chemistry)
1	Remembering/ Recall	15%	15%
2	Understanding/ Comprehension	35%	40%
3	Applying and analysing	40%	35%
4	Evaluating and creating (synthesis)	10%	10%

2.5 Weighting of prescribed content

Paper 1: Physics Focus							
Content	Marks	Total	Duration	Weighting of cognitive levels			
Mechanics	75	150 marks	2 hours	15	35	40	10
Waves, sound and light	40						
Electricity and magnetism	35						

Paper 2: Chemistry Focus							
Content	Marks	Total	Duration	Weighting of cognitive levels			
Chemical change	70	150 marks	2 hours	15	40	35	10
Chemical systems	10						
Matter and materials	70						

2.6 Skills in Physical Sciences

- Identify and question phenomena:
  - ○ Formulate an investigative question.
  - ○ List all possible variables.
  - ○ Formulate a testable hypothesis.

- Design/Plan of an investigation:
  - ○ Identify variables (dependent, independent and controlled variables).
  - ○ List appropriate apparatus.
  - ○ Plan the sequence of steps which should include, amongst others:
    - - The need for more than one trial to minimise experimental errors.
    - - Identify safety precautions that need to be taken.
    - - Identify conditions that ensure a fair test.
    - - Set an appropriate control.

- Graphs:
  - o Draw accurate graphs from given data/information.
  - o Interpret graphs.
  - o Draw sketch graphs from given information.

- Results:
  - o Identify patterns/relationships in data.
  - o Interpret results.

- Conclusions:
  - o Draw conclusions from given information, e.g. tables, graphs.
  - o Evaluate the validity of conclusions.

- Calculations:
  - o Solve problems using two or more different calculations (multistep calculations).

- Descriptions:
  - o Explain/Describe/Argue the validity of a statement/event using scientific principles.

2.7 Prior knowledge from Grade 9

All skills and application of knowledge learnt in Grade 8 and 9 are applicable to assessment in Grade 10. Skills and knowledge from Grade 8 and 9 that may be assessed in Grade 10 include the following:

- The scientific method
- Resistors in series and parallel
- Contact and non-contact forces
- Properties of materials
- The periodic table
- Writing of formulae
- Writing of balanced equations
- Reactions of metals and non-metals with oxygen
- Reactions of metals with acids

3. ELABORATION OF THE CONTENT FOR GRADE 10 (CAPS)

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

3.1 Paper 1: Physics

Vectors and scalars
(This section must be read in conjunction with the CAPS, p. 53.)

Introduction to vectors & scalars
- List physical quantities, for example time, mass, weight, force, charge, etc.   - Define a vector and a scalar quantity. A vector is a physical quantity with magnitude and direction. A scalar is a physical quantity with magnitude only.   - Represent vectors graphically with an arrow. The length of the arrow represents the magnitude and the arrowhead the direction of the vector.   - Use the force vector as an example to show equality of vectors, negative vectors and addition of vectors in one dimension only.   - Define a resultant as the single vector having the same effect as two or more vectors together.   - Determine a resultant graphically using the tail-to-head method as well as by calculation for a maximum of four force vectors in one dimension only, i.e. along a straight line.

Motion in one dimension
(This section must be read in conjunction with the CAPS, p. 54–55.)

Reference frame, position, displacement and distance
- Describe the concept of a frame of reference as a coordinate system used to represent and measure properties of objects, such as position.
- Explain that a frame of reference has an origin and a set of directions, e.g. east and west or up and down.
- Define one-dimensional motion as motion along a straight line. The object may move forward or backward along this line.
- Define position relative to a reference point and understand that position can be positive or negative.
- Define distance as the total path length travelled. Know that distance is a scalar quantity.
- Define displacement as the difference in position in space. Know that displacement is a vector quantity that points from the initial to the final position.
- Describe and illustrate the difference between displacement and distance.
- Calculate distance and displacement for one-dimensional motion.

Average speed, average velocity, acceleration
- Define average speed as the total distance travelled per total time. Know that average speed is a scalar quantity.
- Define average velocity as the rate of change of position.

  In symbols: $v = \dfrac{\Delta x}{\Delta t}$

  Know that average velocity is a vector quantity.
- Calculate average speed and average velocity for one-dimensional motion.
- Convert between different units of speed and velocity, e.g. $m \cdot s^{-1}$ and $km \cdot h^{-1}$.

- Define acceleration as the rate of change of velocity.

  In symbols: $a = \dfrac{\Delta v}{\Delta t}$

- Know that acceleration is a vector quantity. Differentiate between positive acceleration, negative acceleration and deceleration.

  Positive acceleration: An object moving in the positive direction is experiencing an increase in speed and an object moving in the negative direction is experiencing a decrease in speed.

  Negative acceleration: An object moving in the positive direction is experiencing a decrease in speed and an object moving in the negative direction is experiencing an increase in speed

  Deceleration: An object is experiencing a decrease in speed.

Instantaneous speed and velocity and the equations of motion
(This section must be read in conjunction with the CAPS, p. 56–57.)

Instantaneous velocity and instantaneous speed
- Define instantaneous velocity as the rate of change in position, i.e. the displacement divided by a very small time interval or the velocity at a particular time. Know that instantaneous velocity is a vector quantity.
- Define instantaneous speed as the magnitude of the instantaneous velocity. Know that instantaneous speed is a scalar quantity.

Description of motion in words, diagrams, graphs and equations
- Describe in words and distinguish between motion with uniform velocity and uniformly accelerated motion.

  Uniform velocity: Motion at constant velocity, i.e. no acceleration

  Uniform accelerated motion: The velocity of an object changes with the same amount during each time interval.
- Describe the motion of an object given its position versus time, velocity versus time and acceleration versus time graph.
- Determine the velocity of an object from the gradient of the position versus time graph.
- Determine the instantaneous velocity at a particular time using the gradient of a tangent to a position versus time graph.
- Determine the acceleration of an object from the gradient of the velocity vs. time graph.
- Determine the displacement of an object by finding the area between the time axis and the graph of a velocity vs. time graph.
- Use the equations of motion, listed below, to solve problems involving motion in one dimension in the horizontal plane only.

  $v_f = v_i + a\Delta t$

  $\Delta x = v_i \Delta t + \frac{1}{2} a \Delta t^2$

  $v_f^2 = v_i^2 + 2a\Delta x$

  $\Delta x = \left(\dfrac{v_i + v_f}{2} \right) \Delta t$

- Solve problems for the motion of a vehicle including safety issues such as the relationship between speed and stopping distance.

Energy
(This section must be read in conjunction with the CAPS, p. 58–59.)

Gravitational potential energy
- Define gravitational potential energy of an object as the energy it has because of its position in the gravitational field relative to some reference point.
- Calculate the gravitational potential energy of an object using $E_p = mgh$ OR $U = mgh$.

Kinetic energy
- Define kinetic energy as the energy an object possesses as a result of its motion.
- Calculate the kinetic energy of an object using $E_k = \frac{1}{2}mv^2$ OR $K = \frac{1}{2}mv^2$.

Mechanical energy
- Define mechanical energy as the sum of the gravitational potential energy and kinetic energy.
- Calculate mechanical energy using $E_M = E_k + E_p$.   OR $E_M = K + U$

Conservation of mechanical energy
- State the law of the conservation of energy: The total energy of an isolated system remains constant.
  Isolated system: A system that does not interact with its surroundings, i.e. there is no transfer of energy or mass between the system and the surroundings.
- State the principle of conservation of mechanical energy: The total mechanical energy in an isolated system/in the absence of dissipative forces, e.g. friction, remains constant.
  In symbols: $E_{k1} + E_{p1} = E_{k2} + E_{p2}$
- Apply the principle of conservation of mechanical energy to various contexts, viz. objects dropped or thrown vertically upwards, the motion of a pendulum bob, roller coasters and inclined plane problems. In the absence of friction, the mechanical energy of an object moving in the earth's gravitational field is constant (conserved).

Transverse pulses on a string or spring
(This section must be read in conjunction with the CAPS, p. 26.)

Pulse, amplitude
- Define a pulse as a single disturbance in a medium.
- Define a transverse pulse as a pulse in which the particles of the medium move at right angles to the direction of motion of the pulse.
- Define amplitude as the maximum disturbance of a particle from its rest (equilibrium) position.

Superposition of pulses
- Define the principle of superposition as the algebraic sum of the amplitudes of two pulses that occupy the same space at the same time.
- Define constructive interference as the phenomenon where the crest of one pulse overlaps with the crest of another to produce a pulse of increased amplitude.
- Define destructive interference as the phenomenon where the crest of one pulse overlaps with the trough of another, resulting in a pulse of reduced amplitude.
- Apply the principle of superposition to pulses to explain, using diagrams, how two pulses that reach the same point in the same medium superpose constructively and destructively and then continue in the original direction of motion.

Transverse waves
(This section must be read in conjunction with the CAPS, p. 27.)

Wavelength, frequency, amplitude, period, wave speed
•   Define a transverse wave as a wave in which the particles of the medium vibrate at right angles to the direction of motion of the wave. A transverse wave is a succession of transverse pulses.

- Define the terms wavelength, frequency, period, amplitude, crest and trough of a wave.
  Wavelength: The distance between two successive points in phase.
  Frequency: The number of wave pulses per second.
  Period: The time taken for one complete wave pulse.
  Amplitude: The maximum displacement of a particle from its equilibrium position.
  Crest: Highest point (peak) on a wave.
  Trough: Lowest point on a wave.
- Explain the wave concepts in phase and out of phase.
  In phase: Two points in phase are separated by a whole number (1; 2; 3; …) multiple of complete wavelengths.
  Out of phase: Points that are not separated by a whole number multiple of complete wavelengths.
- Identify the wavelength, amplitude, crests, troughs, points in phase and points out of phase on a drawing of a transverse wave.
- Use the relationship between frequency and period, i.e. $f = \dfrac{1}{T}$ and $T = \dfrac{1}{f}$, to solve problems.
- Define wave speed as the distance travelled by a point on a wave per unit time.
- Use the wave equation $v = f\lambda$ to solve problems involving waves.

Longitudinal waves
(This section must be read in conjunction with the CAPS, p. 27-28.)

On a spring
- Define a longitudinal wave as a wave in which the particles of the medium vibrate parallel to the direction of motion of the wave.
- Draw a diagram to represent a longitudinal wave in a spring, showing the direction of motion of the wave relative to the direction in which the particles move.

Wavelength, frequency, amplitude, period, wave speed
- Define the wavelength and amplitude of a longitudinal wave.
  Wavelength: The distance between two successive points in phase.
  Amplitude: The maximum displacement of a particle from its equilibrium position.
- Define a compression as a region of high pressure in a longitudinal wave.
- Define a rarefaction as a region of low pressure in a longitudinal wave.
- Differentiate between longitudinal and transverse waves.
- Define the period and frequency of a longitudinal wave.
  Frequency: The number of wave pulses per second.
  Period: The time taken for one complete wave pulse.
- Use the relationship between frequency and period, i.e. $f = \dfrac{1}{T}$ and $T = \dfrac{1}{f}$, to solve problems.
- Use the wave equation $v = f\lambda$ to solve problems involving longitudinal waves.

Sound
(This section must be read in conjunction with the CAPS, p. 28–29.)

Sound waves
- Explain that sound waves are created by vibrations in a medium in the direction of propagation. The vibrations cause a regular variation in pressure in the medium.
- Describe a sound wave as a longitudinal wave.
- Explain the relationship between wave speed and the properties of the medium in which the wave travels (gas, liquid or solid).
- Describe echoes as reflections of sound waves.
- Use the wave equation $v = f\lambda$ to solve problems involving sound waves that also include echoes, e.g. sonar, bats and dolphins.

Pitch, loudness, quality (tone)
- Relate the pitch of a sound to the frequency of a sound wave. Pitch is the effect produced in the ear due to the sound of a particular frequency. Pitch is directly proportional to frequency.
- Relate the loudness of a sound to both the amplitude of a sound wave and the sensitivity of the human ear. Loudness is a subjective term describing the strength of the ear's perception of a sound. Loudness is directly proportional to amplitude.
- Relate quality of sound to the waveform as it appears to the listener. Two notes of the same pitch and loudness, played on different instruments do not sound the same because the waveforms are different and therefore differ in quality or tone.
- Distinguish between the shape of a pure note and the shape of a noise.

Ultrasound
- Describe sound with frequencies higher than 20 kHz up to about 100 kHz as ultrasound
- Explain how an image can be created using ultrasound based on the fact that when a wave encounters a boundary between two media, part of the wave is reflected, part is absorbed and part is transmitted.
- Describe some of the medical benefits and uses of ultrasound, e.g. safety, diagnosis, treatment and pregnancy.

Electromagnetic radiation
(This section must be read in conjunction with the CAPS, p. 29–31.)

Dual (particle/wave) nature of electromagnetic radiation
Explain that some aspects of the behaviour of electromagnetic radiation can best be explained using a wave model and some aspects can best be explained using a particle model.

Nature of electromagnetic radiation
- Describe the source of electromagnetic waves as an accelerating charge.
- Describe how an electromagnetic wave propagates when an electric field oscillating in one plane produces a magnetic field oscillating in a plane at right angles to it, which produces an oscillating electric field, and so on.
- State that these mutually regenerating fields travel through space at a constant speed of 3×10^8 m·s^{-1}, represented by c.
- List properties of electromagnetic waves:
  - Originate from accelerating electric charges
  - Propagate as electric and magnetic fields that are perpendicular to each other
  - Can travel through a vacuum
  - Have a speed of 3×10^8 m·s^{-1}

Electromagnetic spectrum
- Given a list of different types of electromagnetic radiation, arrange them in order of frequency or wavelength.
- Given the wavelength of electromagnetic waves, calculate the frequency and vice versa, using the equation $c = f\lambda$.
- Give an example of the use of each type of electromagnetic radiation, i.e. gamma rays, X-rays, ultraviolet light, visible light, infrared, microwave and radio and TV waves.
- Indicate the penetrating ability of the different kinds of electromagnetic radiation and relate it to energy of the radiation.
- Describe the dangers of gamma rays, X-rays and the damaging effect of ultra-violet radiation on the skin.

Nature of electromagnetic as particle
- Define a photon as a packet of energy found in light.
- Relate the energy of a photon to the frequency and wavelength of the light.
- Calculate the energy of a photon using $E = hf = \dfrac{hc}{\lambda}$ where $h = 6,63 \times 10^{-34}$ J·s is Planck's constant, $c = 3 \times 10^8$ m·s^{-1} is the speed of light in a vacuum and λ is the wavelength.

Magnetism
(This section must be read in conjunction with the CAPS, p. 38–39.)

Magnetic field of permanent magnets
- Explain that a magnetic field is a region in space where a magnet or ferromagnetic material will experience a force (non-contact).
  Ferromagnetic materials: Materials that are strongly attracted by magnets and are easily magnetised. Examples are iron, cobalt, nickel and their alloys.
  Non-contact force: A force exerted on an object without touching the object.
- Compare magnetic fields with electric and gravitational fields. An electric field is a region in space where an electric charge will experience an electric force. A gravitational field is a region in space where a mass will experience a gravitational force.

Poles of permanent magnets, attraction and repulsion, magnetic field lines
- Describe a magnet as an object that has a pair of opposite poles, called north and south (or north-seeking and south-seeking). Even if the object is cut into tiny pieces, each piece will still have both a north and a south pole.
- Apply the fact that like magnetic poles repel and opposite poles attract to predict the behaviour of magnets when they are brought close together.
- Sketch magnetic field lines to show the shape, size and direction of the magnetic field of different arrangements of bar magnets.
- Describe properties of magnetic field lines:
  o The more closely spaced the field lines are at a point the greater the field at that point.
  o Arrows drawn on the field lines indicate the direction of the field.
  o The direction of a magnetic field points from the North to the South Pole.
  o Magnetic field lines never cross.

Earth's magnetic field, compass
- Explain how a compass indicates the direction of a magnetic field.
- Compare the magnetic field of the Earth to the magnetic field of a bar magnet.
- Explain the difference between the geographical North Pole and the magnetic north pole of the Earth.
  Geographic north pole: Point in the northern hemisphere where the rotation axis of the Earth meets the surface.
  Magnetic north pole: The point where the magnetic field lines of the Earth enters the Earth. It is the direction in which the north pole of a compass points.
  Magnetic south pole: The point where the magnetic field lines of the Earth leaves the Earth.
- Give examples of phenomena that are affected by Earth's magnetic field, e.g. Aurora Borealis (Northern Lights) and magnetic storms.
  Aurora Borealis (Northern Lights): An atmospheric phenomenon consisting of bands of light at the north pole caused by charged solar particles following the Earth's magnetic lines of force.
  Magnetic storm: A disturbance in the Earth's outer magnetosphere, usually caused by streams of charged particles given off by solar flares.
  Magnetosphere: A region surrounding the Earth (extending from about one hundred to several thousand kilometres above the surface) in which charged particles are trapped and their behaviour is dominated by the Earth's magnetic field.
- Discuss qualitatively how the Earth's magnetic field provides protection from solar winds.
  Solar wind: A stream of radioactive and charged particles sent into space at high speeds due to reactions on the sun.

Electrostatics
(This section must be read in conjunction with the CAPS, p. 40–42.)

Two kinds of charge
- State that:
  o All materials contain positive charges (protons) and negative charges (electrons)
  o An object that has an equal number of electrons and protons is neutral (no net charge)
  o Positively charged objects are electron deficient and negatively charged objects have an excess of electrons
- Describe how objects (insulators) can be charged by contact (or rubbing) - tribo-electric charging.
  Tribo-electric charging: A type of contact electrification in which certain materials become electrically charged after they come into contact with different materials and are then separated (such as through rubbing). The polarity and strength of the charges produced differ according to the materials.

Charge conservation
- State that the SI unit for electric charge is the coulomb (C).
- State the principle of conservation of charge: The net charge of an isolated system remains constant during any physical process e.g. two charges making contact and then separating.
- Apply the principle of conservation of charge.
  When two identical conducting objects having charges Q_1 and Q_2 on insulating stands touch, each object has the same final charge on separation.

  Final charge after separation: $Q = \dfrac{Q_1 + Q_2}{2}$

  NOTE: This equation is only true for identically sized conductors on insulated stands.

Charge quantization
- State the principle of charge quantization: All charges in the universe consist of an integer multiple of the charge on one electron, i.e. $1,6 \times 10^{-19}$ C.
- Apply the principle of charge quantization: $Q = nq_e$, where $q_e = 1,6 \times 10^{-19}$ C and n is an integer.

Force exerted by charges on each other (descriptive)
- State that like charges repel and opposite charges attract.
- Explain how charged objects can attract uncharged insulators because of the polarisation of molecules inside insulators.
  Polarisation: The partial or complete polar separation of positive and negative electric charge in a system.

Electric Circuits
(This section must be read in conjunction with the CAPS, p. 42–45.)

Terminal potential difference and emf
- Define potential difference across the ends of a conductor as the energy transferred per unit electric charge flowing through it. In symbols: $V = \dfrac{W}{Q}$

  Potential difference is measured in volts (V).
- Define emf as the work done per unit charge by the source (battery). It is equal to the potential difference measured across the terminals of a battery when no charges are flowing in the circuit.
- Define terminal potential difference as the voltage measured across the terminals of a battery when charges are flowing in the circuit.
- Do calculations using $V = \dfrac{W}{Q}$.

Current
- Define current strength, I, as the rate of flow of charge. It is measured in ampere (A), which is the same as coulomb per second.
- Calculate current strength in a conductor using the equation $I = \dfrac{Q}{\Delta t}$.

  Q is the symbol for electric charge measured in coulomb (C). One coulomb is defined as the charge transferred in a conductor in one second if the current is one ampere.
- Indicate the direction of conventional current (from positive to negative) in circuit diagrams using arrows.

Measurement of potential difference and current
- Draw a diagram to show how to correctly connect an ammeter to measure the current through a given circuit element. An ammeter is connected in series and has a very low resistance.
- Draw a diagram to show how to correctly connect a voltmeter to measure the potential difference across a given circuit element. A voltmeter is connected in parallel and has a very high resistance.

Resistance
- Define resistance as the ratio of the potential difference across a resistor to the current in the resistor.
- Explain that resistance is the opposition to the flow of electric charges.
- Define the unit of resistance: One ohm (Ω) is equal to one volt per ampere.
- Give a microscopic description of resistance in terms of electrons moving through a conductor colliding with the particles of which the conductor (metal) is made and transferring kinetic energy.
- State and explain factors that affect the resistance of a given material, i.e. temperature, length and thickness.
- Explain why a battery in a circuit goes flat eventually by referring to the energy transformations that take place in the battery and the resistors in a circuit.

Resistors in series
- Know that current is the same through each resistor in a series circuit.
- Describe series circuits as potential difference dividers because the total potential difference is equal to the sum of the potential differences across all the individual components.
- Calculate the equivalent (total) resistance of resistors connected in series using $R_s = R_1 + R_2 + \ldots$

Resistors in parallel
- Know that potential difference is the same across resistors connected in parallel.
- Describe parallel circuits as current dividers because the total current in the circuit is equal to the sum of the branch currents.
- Calculate the equivalent (total) resistance of resistors connected in parallel using
$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} + \ldots$$

3.2   Paper 2: Chemistry

Matter and classification
(This section must be read in conjunction with the CAPS, p. 15–18.)

The material(s) of which an object is composed
- Describe matter as being made up of particles whose properties determine the observable characteristics of matter and its reactivity.
- Define properties of materials:
  - Strength
  - Brittle: Hard but likely to break easy.
    Malleable: Ability to be hammered or pressed into shape without breaking or cracking.
    Ductile: Ability to be stretched into a wire.
  - Density: The mass per unit volume of a substance.
  - Melting points and boiling points
    Boiling point: The temperature of a liquid at which its vapour pressure equals the external (atmospheric) pressure.
    Melting point: The temperature at which a solid, given sufficient heat, becomes a liquid.

Mixtures: heterogeneous and homogeneous
- Define a homogeneous mixture as a mixture of uniform composition and in which all components are in the same phase, e.g. a solution of salt and water.
- Define a heterogeneous mixture as a mixture of non-uniform composition and of which the components can be easily identified, e.g. sand and water.
- Give examples of heterogeneous and homogeneous mixtures.
- Classify given mixtures as homogenous and heterogeneous.

Pure substances: elements and compounds
- Use symbols to represent elements and compounds.
- Define an element as a pure substance consisting of one type of atom.
- Define a compound as a pure substance consisting of two or more different elements.
- Define a pure substance as a substance that cannot be separated into simpler components by physical methods.
- Classify given substances as pure or impure and as compounds or elements.

Names and formulae of substances
- Write names of compounds from given formulae or write down formulae of compounds from given names.
- Write names of ions from given formulae or formulae from given names.
- Write names of substances or ions ending on -ide, -ite and –ate.
- Write names of substances using the prefixes di-, tri-, etc.

Metals, metalloids and non-metals
- Classify substances as metals, metalloids and non-metals using their properties.
- Identify the metals, their positions on the periodic table and their numbers in comparison with the number of non-metals.
- Identify the non-metals and their positions on the periodic table.
- Describe metalloids as having properties of metals and non-metals.
- Describe the characteristic property of metalloids that show increasing conductivity with increasing temperature (the reverse of metals), e.g. silicon and graphite.
- Identify the metalloids and their position on the periodic table.

Electrical conductors, semiconductors and insulators
- Define the terms electrical conductor, semiconductor and electrical insulator:
  - Electrical conductor: A material that allows the flow of charge.
  - Semiconductor: A substance that can conduct electricity under some conditions, but not others, making it a good medium for the control of electrical current.
  - Electrical insulator: A material that prevents the flow of charge.
- Classify materials as electrical conductors, semiconductors and insulators.
- Give examples of electrical conductors, semiconductors and insulators.

Thermal conductors and insulators
- Define the terms thermal conductor and thermal insulator. A thermal conductor is a material that allows heat to pass through easily, whilst a thermal insulator does not allow heat to pass through it.
- Describe a test to classify materials as thermal conductors and insulators.
- Give examples of materials that are thermal conductors and insulators.

Magnetic and nonmagnetic materials
- Describe how to test and classify materials as magnetic and non-magnetic.
- Give examples of materials that are magnetic and non-magnetic.
- Give examples of how we use magnets in daily life (in speakers, telephones, electric motors and as compasses).

States of Matter and the Kinetic Molecular Theory
(This section must be read in conjunction with the CAPS, p. 19.)

Three states of matter
- Describe the particle nature of matter by referring to diffusion and Brownian motion.
  Diffusion: The movement of atoms or molecules from an area of higher concentration to an area of lower concentration.
  Brownian motion: The random movement of microscopic particles suspended in a liquid or gas, caused by collisions between these particles and the molecules of the liquid or gas.
- List and characterise the three states of matter.
- Define freezing point, melting point and boiling point.
  Boiling point: The temperature of a liquid at which its vapour pressure equals the external (atmospheric) pressure.
  Melting point: The temperature at which a solid, given sufficient heat, becomes a liquid.
  Freezing point: The temperature at which a liquid changes to a solid by the removal of heat.
- Interpret/Draw heating and cooling curves and interpret data given on heating and cooling curves.
- Identify the physical state of a substance at a specific temperature, given the melting point and the boiling point of the substance.
- Define melting, evaporation, freezing, sublimation and condensation as changes in state.
  Melting: The process during which a solid changes to a liquid by the application of heat.
  Evaporation: The change of a liquid into a vapour at any temperature below the boiling point. (Note: Evaporation takes place at the surface of a liquid, where molecules with the highest kinetic energy are able to escape. When this happens, the average kinetic energy of the liquid is lowered, and its temperature decreases.

Freezing: The process during which a liquid changes to a solid by the removal of heat.

Sublimation: The process during which a solid changes directly into a gas without passing through an intermediate liquid phase.

Condensation: The process during which a gas or vapour changes to a liquid, either by cooling or by being subjected to increased pressure.

Kinetic Molecular Theory
- Describe a solid, a liquid, and a gas according to the Kinetic Molecular Theory in terms of particles of matter. According to the Kinetic Molecular Theory:
  o Matter consists of small particles
  o The particles are in constant motion
  o There are forces of attraction between the particles
  o Particles collide (with the sides of the container and each other) and exert pressure
  o The temperature of a substance is a measure of the average kinetic energy of the particles
  o A phase change may occur when the energy of particles changes

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

Models of the atom
- Describe the major contributions (Dalton, Thomson, Rutherford, Bohr, Chadwick) to the atomic model used today.

Structure of the atom: protons, neutrons, electrons
- Define the atomic number as the number of protons in an atom of an element.
- Given a periodic table or suitable data, determine for an atom/ion the:
  o Atomic number
  o Number of protons
  o Number of electrons
  o Number of neutrons
  o Mass number
- Show that by removing electrons from an atom the neutrality of the atom is changed.
- Determine the charge on an ion after removing electrons from or adding electrons to an atom.

Isotope
- Define isotopes as atoms of the same element having the same number of protons, but different numbers of neutrons.
- Define relative atomic mass as the mass of a particle on a scale where an atom of carbon-12 has a mass of 12.
- Calculate the relative atomic mass of naturally occurring elements from the percentage of each isotope in a sample of the naturally occurring element and the relative atomic mass of each of the isotopes.
- Represent atoms using the notation $^A_Z E$ where E is the symbol of the element, Z is the atomic number and A is the mass number.

Electron configuration
- Use Aufbau diagrams (orbital box diagrams) and the electron configuration notation (sp notation) to give electronic arrangements of atoms up to Z = 20.
- Know that every orbital corresponds to a specific energy value that electrons have when occupying it. Describe atomic orbitals as the most probable regions in space where electrons that have the specific energy corresponding to the orbital are found.

- Describe the shape of s-orbitals as spherical and that of p-orbitals as pairs of dumb-bells aligned along the x-, y- and z-axes at 90° to each other.
- State Hund's rule: No pairing in p orbitals before there is not at least one electron in each of them.
- State Pauli's Exclusion Principle: Maximum of two electrons per orbital provided that they spin in opposite directions.

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

The positions of the elements in the periodic table related to their electronic arrangements
- Describe the periodic table as displaying the elements in order of increasing atomic number and showing how periodicity of the physical and chemical properties of the elements relates to atomic structure.
- Define the group number and the period number of an element in the periodic table.
  Groups are the vertical columns in the periodic table. Some groups have names, e.g. alkali metals (group I), earth-alkaline metals (group II), halogens (group 17 or VII) and noble gases (group18 or VIII).
  Periods are the horizontal rows in the periodic table.
- Relate the position of an element in the periodic table to its electronic structure and vice versa.
- Describe periodicity from Li to Ar in terms of atomic radius, ionisation energy, electron-affinity and electronegativity. Describe the changes in terms of change in charge of the nucleus and distance between the nucleus and the electron. Periodicity is the repetition of similar properties in chemical elements, as indicated by their positioning in the periodic table.
- Define atomic radius, ionisation energy, electron-affinity and electronegativity.
  Atomic radius: Radius of an atom, i.e. the mean distance from the nucleus to the border of the outer orbital.
  Ionisation energy: Energy needed per mole to remove an electron(s) from an atom in the gaseous phase.
  First ionisation energy: Energy needed per mole to remove the first electron from an atom in the gaseous phase.
  Electron affinity: The energy released when an electron is attached to an atom or molecule to form a negative ion.
  Electronegativity: A measure of the tendency of an atom in a molecule to attract bonding electrons.

Similarities in chemical properties among elements in Groups 1, 2, 17 and 18
- Relate the electronic arrangements to the chemical properties of group 1, 2, 17 and 18 elements.
- Describe the trend in reactivity of elements in groups 1, 2 and 17.
  Groups 1 and 2: Chemical reactivity increases from top to bottom.
  Group 17: Chemical reactivity decreases from top to bottom.
- Predict chemical properties of unfamiliar elements in groups 1, 2, 17 and 18 of the periodic table.
- Indicate that metals are found on the left-hand side of the periodic table.
- Indicate that non-metals are found on the right-hand side of the periodic table.
- Indicate where transition metals are to be found on the periodic table.

Chemical bonding
(This section must be read in conjunction with the CAPS, p. 25.)

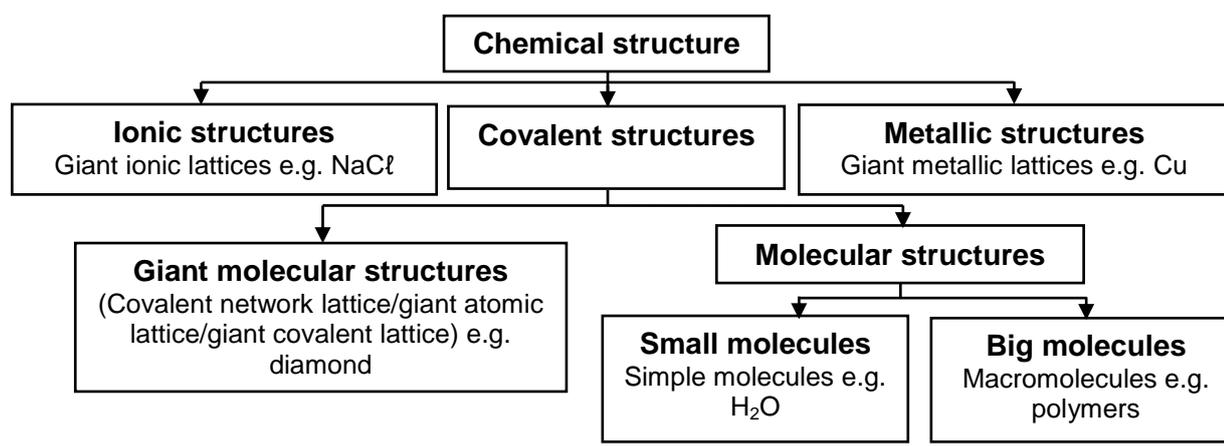**Covalent bonding, ionic bonding and metallic bonding**

- Define a chemical bond as a mutual attraction between two atoms resulting from the simultaneous attraction between their nuclei and the outer electrons. (The energy of the combined atoms is lower than that of the individual atoms resulting in higher stability.)
- Draw Lewis dot diagrams of elements.
  A Lewis dot diagram is a structural formula in which valence electrons are represented by dots or crosses. It is also known as an electron dot formula, a Lewis formula, or an electron diagram.
- Define a covalent bond as the sharing of electrons between atoms to form molecules.
  Molecule: A group of two or more atoms that are covalently bonded and that functions as a unit.
- Draw Lewis dot diagrams of simple covalent molecules containing single, double and triple covalent bonds: H_2; F_2, $C\ell_2$, O_2, N_2, HF, $HC\ell$, CH_4, NH_3, H_2O
  In a Lewis dot diagram two dots between atoms represent a covalent bond. These two electrons are known as a bonding pair, whilst non-binding electron pairs are called lone pairs.
- Write names and formulae of covalent compounds.
- Define ionic bonding as the transfer of electrons to form cations (positive ions) and anions (negative ions) that attract each other to form a formula-unit.
  A formula-unit is the most simple empirical formula that represents the compound.
  An ion is a charged particle made from an atom by the loss or gain of electrons.
  An anion (negative ion) is a charged particle made from an atom by the gain of electrons
  A cation (positive ion) is a charged particle made from an atom by the loss of electrons.
- Draw Lewis dot diagrams of cations and anions.
- Draw Lewis dot diagrams to show the formation of simple ionic compounds such as $NaC\ell$, $KC\ell$, KBr, $CaC\ell_2$ and $MgBr_2$.
- Predict the ions formed by atoms of metals and non-metals by using information in the periodic table. Metals occur on the left-hand side of the periodic table and form positive ions, whilst non-metals occur on the right-hand side of the periodic table and form negative ions.
- Name ionic compounds based on the component ions.
- Describe the structure of the sodium chloride crystal. In the crystal each sodium ion is surrounded by six chloride ions to form a cubic structure. Each chloride ion is also surrounded by six sodium ions.
  A crystal lattice: An orderly three-dimensional arrangement of particles (ions, molecules or atoms) in a solid structure.
- Define metallic bonding as the bond between positive ions and delocalised valence electrons in a metal.
  Valence electrons or outer electrons are the electrons in the highest energy level of an atom in which there are electrons.
- Calculate relative molecular masses for covalent molecules, e.g. $M_r(HC\ell) = 35,5$.
- Calculate relative formula masses for ionic compounds, e.g. $M_r(NaC\ell) = 57,5$.

Particles substances are made of
(This section must be read in conjunction with the CAPS, p. 32–34.)

Atoms and compounds
- Describe atoms as the smallest particles of which all substances are made.
- Identify noble gases as the only substances found in atomic form at ambient conditions.
- Describe a compound as a group of two or more different atoms that attract each other by relatively strong forces or bonds. The atoms combine in definite proportions.
- Classify substances as covalent, ionic or metallic structures.
- Describe covalent structures as consisting of molecular structures and giant molecular structures.
- Describe molecular structures. They:
  - Consist of molecules formed when atoms share electrons
  - Form due to intermolecular forces between molecules, e.g. oxygen, water, petrol, CO_2, S_8, C_{60} (buckminsterfullerene or buckyballs), consisting of non-metallic atoms which are covalently bonded
  - Are represented using molecular formula, e.g. H_2O, C_8H_{18}
- Describe giant molecular structures or covalent network structures. They:
  - Consist of atoms
  - Form due to covalent bonds (sharing of electrons) between atoms to form giant repeating lattices of covalently bonded atoms
  - Are e.g. diamond and graphite, both with empirical formula C and quartz, glass or sand, all with empirical formula SiO_2
- Describe ionic structures. They:
  - Consist of positive and negative ions
  - Are formed when the electrons of atoms are transferred from one atom to another atom to form positive and negative ions
  - Are due to ionic bonding between positive and negative ions
  - Are usually composed of both metallic elements (usually forming positive ions) and non-metallic elements (usually forming negative ions)
  - Are also called ionic substances or ionic solids or ionic compounds, e.g. a sodium chloride crystal and a potassium permanganate crystal
- Describe metallic structures. They:
  - Consist of positive ions and delocalised valence electrons
  - Are formed when metal atoms lose their outer electrons (valence electrons) to form a lattice of regularly spaced positive ions and the outer electrons form a delocalised pool of electrons that surround the positive ions
  - Are due to metallic bonding between positive ions and delocalised valance electrons
  - Are e.g. metal crystals like a piece of copper, or zinc, or iron
- Classify all chemical structures as one of ionic structures, giant molecular structures, molecular structures or metallic structures, as illustrated in the flow diagram below.

Chemical structure

Ionic structures
Giant ionic lattices e.g. NaCℓ

Covalent structures

Metallic structures
Giant metallic lattices e.g. Cu

Giant molecular structures
(Covalent network lattice/giant atomic lattice/giant covalent lattice) e.g. diamond

Molecular structures

Small molecules
Simple molecules e.g. H_2O

Big molecules
Macromolecules e.g. polymers

Physical and Chemical Change
(This section must be read in conjunction with the CAPS, p. 35.)

Separation of particles in physical and chemical change
- Define a physical change as a change in which:
    - No new substances are formed
    - Energy changes are small in relation to chemical changes
    - Mass, numbers of atoms and molecules as being conserved
- Describe the rearrangement of molecules during physical changes, e.g.
    - Molecules separate when water evaporates to form water vapour
    - When ice melts molecules become disorderly arranged due to breaking of intermolecular forces
- Define a chemical change as a change in which:
    - New chemical substances are formed
    - Energy changes are much larger than those of the physical change
      Endothermic reaction: Energy is absorbed during the reaction
      Exothermic reaction: Energy is released during the reaction
    - Mass and atoms are conserved, but the number of molecules is not
- Describe examples of a chemical change that include the:
    - Decomposition of hydrogen peroxide to form water and oxygen
    - Synthesis reaction that occurs when hydrogen burns in oxygen to form water
    - Heating of iron and sulphur
    - Reaction of lead(II) nitrate and potassium iodide (in solid phase and/or as solutions)
    - Titration of hydrochloric acid with sodium hydroxide to measure the change in temperature

Conservation of atoms and mass
- Calculate relative molecular masses of reactants and products in balanced equations to illustrate that atoms are conserved during chemical reactions, but not molecules.

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

Balanced chemical equations
- Write and balance chemical equations. Use formulae with subscripts to represent phases, viz. (s), (ℓ), (g) and (aq).
- Interpret balanced reaction equations in terms of:
    - Conservation of atoms
    - Conservation of mass (use relative atomic masses)

Reactions in aqueous solutions
(This section must be read in conjunction with the CAPS, p. 46–49.)

Ions in aqueous solution: their interaction and effects
- Explain, using diagrams, the polar nature of the water molecule and how water is able to dissolve ions.
- Define a polar molecule as having two oppositely charged poles and that it is also known as a dipole.
- Represent the dissolution process using balanced reaction equations with the abbreviation (s) for the solid phase and (aq) for substances dissolved in water, e.g. when salt is dissolved in water ions form according to the equation:
  $NaCℓ(s) \rightarrow Na^+(aq) + Cℓ^-(aq)$
- Define an aqueous solution as a solution in which the solvent is water.
- Define dissociation as the process in which solid ionic crystals are broken up into ions when dissolved in water.

- Define terms such as:

  Hydration: The process in which ions are surrounded with water molecules

  Solubility: The maximum amount of a substance (the solute) that may be dissolved in another (the solvent).

  Solute: The dissolved substance in a solution – usually the substance present in lesser amount

  Solution: A homogenous mixture of two or more substances.

  Solvent: The substance in a solution in which the solute is dissolved - usually the substance present in greater amount

Electrolytes and extent of ionization as measured by conductivity
- Describe/Interpret a simple circuit to measure conductivity of solutions.
- Define conductivity as the ability of a material to conduct electricity.
- Define an electrolyte as a solution that conducts electricity through the movement of ions.
- Relate conductivity to the:
  o Concentration of ions in solution and this in turn to the solubility of ionic substances
    Define concentration as the amount of substance present per volume of a solution.
  o Type of substance, since some substances, like sugar, dissolve but this does not affect conductivity

Precipitation reactions
- Write balanced equations to describe the precipitation of insoluble salts.
- Explain how to test for the presence of the following anions in solution and write chemical equations:
  o Chlorides – using silver nitrate and nitric acid
  o Bromides – using silver nitrate and nitric acid
  o Iodides – using silver nitrate and nitric acid
  o Sulphates – using barium nitrate and nitric acid
  o Carbonates – using barium nitrate and nitric acid (precipitate dissolves in nitric acid) or acid and calcium hydroxide (clear lime water)
- Identify an ion or ions in a solution from a description of the reactants mixed and the observations of the products.

Other chemical reaction types in water solution
- Classify different ion exchange reactions as:
  o Precipitation reactions – reactions in which an insoluble product forms when solutions are mixed
  o Gas-forming reactions – reactions in which the driving force is the formation of a gas as one of the products
  o Acid-base reactions – reactions in which a hydrogen ion (H^+ ion) is transferred from one of the reactants to another
- Classify reactions as redox reactions, i.e. reactions in which an electron transfer takes place. One reactant gains electrons and another loses electrons. Use the charge of the atom to determine whether electrons are transferred in the following reactions:
  o Reactions of acids ($HC\ell$ & H_2SO_4) with metals to form metallic salts, e.g.
    $2HC\ell(aq) + Zn(s) \rightarrow ZnC\ell_2(aq) + H_2(g)$
  o $Fe(s) + S(s) \rightarrow FeS(s)$
  o $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$
- Use examples of reactions done that are not redox reactions, e.g. precipitation reactions, to show that there is no change in charge of atoms.

Quantitative Aspects of Chemical Change
(This section must be read in conjunction with the CAPS, p. 50–52.)

Atomic mass and the mole concept
- Describe the mole as the SI unit for amount of substance.
- Define one mole as the amount of substance having the same number of particles as there are atoms in 12 g carbon-12.
- Define relative atomic mass as the mass of a particle on a scale where an atom of carbon-12 has a mass of 12.
- Describe Avogadro's number, N_A, as the number of particles (atoms, molecules, formula-units) present in mole ($N_A = 6,023 \times 10^{23}$ mol^{-1}).
- Define molar mass as the mass of one mole of a substance measured in g·mol^{-1}.
- Describe the relationship between molar mass and relative molecular mass and relative formula mass.
- Calculate the molar mass of a substance given its formula.

Molecular and formula masses

- Calculate mass, molar mass and number of moles according to the relationship $n = \dfrac{m}{M}$.

- Determine the empirical formula for a given substance from percentage composition.
  Define an empirical formula as the simplest whole-number ratio of atoms in a compound.
- Determine the number of moles of water of crystallisation in salts like $A\ell C\ell_3 \cdot nH_2O$.
  Define water of crystallisation as water that is stoichiometrically bound into a crystal, e.g. the H_2O in $CuSO_4 \cdot 5H_2O$.

Determining the composition of substances
- Determine percentage composition of an element in a compound.
  Percentage composition is the mass of each atom present in a compound expressed as a percentage of the total mass of the compound.
- Define concentration as the number of moles of solute per cubic decimetre of solution.

- Calculate concentration in mol·dm^{-3} using $c = \dfrac{n}{V}$.

Molar volume of gases
- State Avogadro's law, i.e. one mole of any gas occupies the same volume at the same temperature and pressure.
- At STP: 1 mole of any gas occupies 22,4 dm^3 at 0 °C (273 K) and 1 atmosphere (101,3 kPa). Thus the molar gas volume, V_M, at STP = 22,4 dm^3·mol^{-1}.
- 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 occupies the same volume.

Basic stoichiometric calculations
- Perform stoichiometric calculations based on balanced equations. These may include calculations based on concentration, mass, moles, molar mass, number of particles and volume.
- Determine the theoretical yield of a product in a chemical reaction when you start with a known mass of reactant. The theoretical yield is the calculated yield of a product in a chemical reaction. Actual yield is the quantity physically obtained from a chemical reaction.
- Determine the percentage yield of a chemical reaction: $\%\text{yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$

The hydrosphere
(This section must be read in conjunction with the CAPS, p. 60.)

Its composition and interaction with other global systems

• Describe the hydrosphere as the water of the Earth and is found as liquid water (surface and underground), ice (polar ice, icebergs and ice frozen in the soil called permafrost) and water vapour in the atmosphere.
 • Describe the interaction of the hydrosphere with the atmosphere, the lithosphere and the biosphere.
     Atmosphere: The body of air surrounding the Earth.
     Lithosphere: The solid, rocky crust covering the entire planet.
     Biosphere: All the living organisms, i.e. plants and animals.
 • Describe the water cycle and interpret diagrams of the water cycle.
 • Explain how the building of dams affects the lives of the people and the ecology in the region.

4.   GENERAL INFORMATION

4.1   Quantities, symbols and units

The most common quantities, symbols and SI units used in introductory Physics are listed below. **A quantity should not be confused with the units in which it is measured**.

Quantity	Preferred symbol	Alternative symbol	Unit name	Unit symbol
mass	m		kilogram	kg
position	x, y		metre	m
displacement	$\Delta x, \Delta y$	s	metre	m
velocity	v_x, v_y	u, v	metre per second	$m \cdot s^{-1}$
initial velocity	v_i	u	metre per second	$m \cdot s^{-1}$
final velocity	v_f	v	metre per second	$m \cdot s^{-1}$
acceleration	a		metre per second per second	$m \cdot s^{-2}$
acceleration due to gravity	g		metre per second per second	$m \cdot s^{-2}$
time (instant)	t		second	s
time interval	Δt		second	s
energy	E		joule	J
kinetic energy	K	E_k	joule	J
potential energy	U	E_p	joule	J
work	W		joule	J
force	F		newton	N
weight	w	F_g	newton	N
wavelength	λ		metre	m
frequency	f	ν	hertz or per second	Hz or s^{-1}
period	T		second	s
speed of light	c		metre per second	$m \cdot s^{-1}$
charge	Q, q		coulomb	C
potential difference	$\Delta V, V$		volt	V
emf	**E**	ε	volt	V
current	I, i		ampere	A
resistance	R		ohm	Ω

Conventions (e.g. signs, symbols, terminology and nomenclature)
The syllabus and question papers will conform to generally accepted international practices.

NOTE:
1.   For marking purposes, alternative symbols will also be accepted.
2.   Separate compound units with a multiplication dot, not a full stop, for example $m \cdot s^{-1}$.
      For marking purposes, $m.s^{-1}$ will also be accepted.
3.   Use the equal sign only when it is mathematically correct, for example:
      Incorrect: 1 cm = 1 m   (on a scale drawing)
      Correct: 1 cm = 10^{-2} m   1 cm represents 1 m (on a scale drawing)

4.2   Information sheets – Paper 1 (Physics)

TABLE 1: PHYSICAL CONSTANTS/*TABEL 1: FISIESE KONSTANTES*

NAME/*NAAM*	SYMBOL/*SIMBOOL*	VALUE/*WAARDE*
Acceleration due to gravity *Swaartekragversnelling*	g	$9,8 \text{ m·s}^{-2}$
Speed of light in a vacuum *Spoed van lig in 'n vakuum*	c	$3,0 \times 10^8 \text{ m·s}^{-1}$
Planck's constant *Planck se konstante*	h	$6,63 \times 10^{-34} \text{ J·s}$
Charge on electron *Lading op elektron*	e	$-1,6 \times 10^{-19} \text{ C}$
Electron mass *Elektronmassa*	m_e	$9,11 \times 10^{-31} \text{ kg}$

TABLE 2: FORMULAE/*TABEL 2: FORMULES*

MOTION/*BEWEGING*

$v_f = v_i + a\Delta t$	$\Delta x = v_i \Delta t + \frac{1}{2} a \Delta t^2$
$v_f^2 = v_i^2 + 2a\Delta x$	$\Delta x = \left(\dfrac{v_f + v_i}{2} \right) \Delta t$

WORK, ENERGY AND POWER/*ARBEID, ENERGIE EN DRYWING*

$U = mgh$ or/*of* $E_P = mgh$	$K = \frac{1}{2}mv^2$ or/*of* $E_k = \frac{1}{2}mv^2$

WAVES, SOUND AND LIGHT/*GOLWE, KLANK EN LIG*

$v = f\lambda$	$T = \dfrac{1}{f}$
$E = hf$ or/*of* $E = h\dfrac{c}{\lambda}$	

ELECTRIC CIRCUITS/*ELEKTRIESE STROOMBANE*

$Q = I \Delta t$	$\dfrac{1}{R_p} = \dfrac{1}{R_1} + \dfrac{1}{R_2} + ...$
$R_s = R_1 + R_2 + ...$	$V = \dfrac{W}{Q}$

4.3   Information sheets – Paper 2 (Chemistry)

TABLE 1: PHYSICAL CONSTANTS/*TABEL 1: FISIESE KONSTANTES*

NAME/*NAAM*	SYMBOL/*SIMBOOL*	VALUE/*WAARDE*
Standard pressure *Standaarddruk*	p^θ	$1{,}013 \times 10^5$ Pa
Molar gas volume at STP *Molêre gasvolume by STD*	V_m	$22{,}4$ dm$^3 \cdot$mol^{-1}
Standard temperature *Standaardtemperatuur*	T^θ	273 K
Charge on electron *Lading op elektron*	e	$-1{,}6 \times 10^{-19}$ C
Avogadro's constant *Avogadro-konstante*	N_A	$6{,}02 \times 10^{23}$ mol^{-1}

TABLE 2: FORMULAE/*TABEL 2: FORMULES*

$n = \dfrac{m}{M}$	$n = \dfrac{N}{N_A}$
$c = \dfrac{n}{V} \quad$ OR $\quad c = \dfrac{m}{MV}$	$n = \dfrac{V}{V_m}$

TABLE 3: THE PERIODIC TABLE OF ELEMENTS/*TABEL 3: DIE PERIODIEKE TABEL VAN ELEMENTE*

KEY/*SLEUTEL*

Atomic number
Atoomgetal

Electronegativity
Elektronegatiwiteit →

$\begin{array}{c} 29 \\ 1,9 \quad \textbf{Cu} \\ 63,5 \end{array}$

← Symbol
Simbool

Approximate relative atomic mass
Benaderde relatiewe atoommassa

1 (I)	2 (II)	3	4	5	6	7	8	9	10	11	12	13 (III)	14 (IV)	15 (V)	16 (VI)	17 (VII)	18 (VIII)
1 **H** 2,1 1																	2 **He** 4
3 **Li** 1,0 7	4 **Be** 1,5 9											5 **B** 2,0 11	6 **C** 2,5 12	7 **N** 3,0 14	8 **O** 3,5 16	9 **F** 4,0 19	10 **Ne** 20
11 **Na** 0,9 23	12 **Mg** 1,2 24											13 **Aℓ** 1,5 27	14 **Si** 1,8 28	15 **P** 2,1 31	16 **S** 2,5 32	17 **Cℓ** 3,0 35,5	18 **Ar** 40
19 **K** 0,8 39	20 **Ca** 1,0 40	21 **Sc** 1,3 45	22 **Ti** 1,5 48	23 **V** 1,6 51	24 **Cr** 1,6 52	25 **Mn** 1,5 55	26 **Fe** 1,8 56	27 **Co** 1,8 59	28 **Ni** 1,8 59	29 **Cu** 1,9 63,5	30 **Zn** 1,6 65	31 **Ga** 1,6 70	32 **Ge** 1,8 73	33 **As** 2,0 75	34 **Se** 2,4 79	35 **Br** 2,8 80	36 **Kr** 84
37 **Rb** 0,8 86	38 **Sr** 1,0 88	39 **Y** 1,2 89	40 **Zr** 1,4 91	41 **Nb** 92	42 **Mo** 1,8 96	43 **Tc** 1,9	44 **Ru** 2,2 101	45 **Rh** 2,2 103	46 **Pd** 2,2 106	47 **Ag** 1,9 108	48 **Cd** 1,7 112	49 **In** 1,7 115	50 **Sn** 1,8 119	51 **Sb** 1,9 122	52 **Te** 2,1 128	53 **I** 2,5 127	54 **Xe** 131
55 **Cs** 0,7 133	56 **Ba** 0,9 137	57 **La** 139	72 **Hf** 1,6 179	73 **Ta** 181	74 **W** 184	75 **Re** 186	76 **Os** 190	77 **Ir** 192	78 **Pt** 195	79 **Au** 197	80 **Hg** 201	81 **Tℓ** 1,8 204	82 **Pb** 1,8 207	83 **Bi** 1,9 209	84 **Po** 2,0	85 **At** 2,5	86 **Rn**
87 **Fr** 0,7	88 **Ra** 0,9 226	89 **Ac**															

58 **Ce** 140	59 **Pr** 141	60 **Nd** 144	61 **Pm**	62 **Sm** 150	63 **Eu** 152	64 **Gd** 157	65 **Tb** 159	66 **Dy** 163	67 **Ho** 165	68 **Er** 167	69 **Tm** 169	70 **Yb** 173	71 **Lu** 175
90 **Th** 232	91 **Pa**	92 **U** 238	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

5.   MARKING GUIDELINES: PAPER 1

5.1   Calculations

5.1.1 Marks will be awarded for: correct formula, correct substitution, correct answer with unit.

5.1.2 No marks will be awarded if an **incorrect or inappropriate formula is used**, even though there may be relevant symbols and applicable substitutions.

5.1.3 When an error is made during **substitution into a correct formula**, a mark will be awarded for the correct formula and for the correct substitutions, but **no further marks** will be given.

5.1.4 If **no formula** is given, but **all substitutions are correct**, the candidate will **forfeit one mark**.

5.1.5 No penalisation if zero substitutions are omitted in calculations where **correct formula/principle is given correctly**.

5.1.6 Mathematical manipulations and change of subject of appropriate formulae carry no marks, but if a candidate starts off with the correct formula and then changes the subject of the formula incorrectly, marks will be awarded for the formula and the correct substitutions. The mark for the incorrect numerical answer is forfeited.

5.1.7 Marks are only awarded for a formula if a **calculation has been attempted**, i.e. substitutions have been made or a numerical answer given.

5.1.8 Marks can only be allocated for substitutions when values are substituted into formulae and not when listed before a calculation starts.

5.1.9 Final answers to all calculations, when not specified in the question, must be rounded off to a minimum of TWO decimal places.

5.1.10 If a final answer to a calculation is correct, full marks will not automatically be awarded. Markers will always ensure that the correct/appropriate formula is used and that workings, including substitutions, are correct.

5.1.11 Questions in which a series of calculations have to be made (e.g. a circuit-diagram question) do not necessarily always have to follow the same order. FULL MARKS will be awarded, provided it is a valid solution to the problem. However, any calculation that will not bring the candidate closer to the answer than the original data, will not count any marks.

5.2   Units

5.2.1 Candidates will only be penalised once for the repeated use of an incorrect unit **within a question**.

5.2.2 Units are only required in the final answer to a calculation.

5.2.3 Marks are only awarded for an answer, and not for a unit per se. Candidates will therefore forfeit the mark allocated for the answer in each of the following situations:
- Correct answer + wrong unit
- Wrong answer + correct unit
- Correct answer + no unit

5.2.4 SI units must be used, except in certain cases, e.g. $V \cdot m^{-1}$ instead of $N \cdot C^{-1}$, and $cm \cdot s^{-1}$ or $km \cdot h^{-1}$ instead of $m \cdot s^{-1}$ where the question warrants this.

5.3 General

5.3.1 If one answer or calculation is required, but two are given by the candidate, only the first one will be marked, irrespective of which one is correct. If two answers are required, only the first two will be marked, etc.

5.3.2 For marking purposes, alternative symbols (s, u, t, etc.) will also be accepted.

5.3.3 Separate compound units with a multiplication dot, not a full stop, e.g. $m \cdot s^{-1}$. For marking purposes, $m.s^{-1}$ and m/s will also be accepted.

5.4 Positive marking

Positive marking regarding calculations will be followed in the following cases:

5.4.1 Subquestion to subquestion: When a certain variable is incorrectly calculated in one subquestion (e.g. 3.1) and needs to be substituted into another subquestion (3.2 or 3.3), **full marks** are to be awarded for the subsequent subquestions.

5.4.2 A multistep question in a subquestion: If the candidate has to calculate, for example, current in the first step and gets it wrong due to a substitution error, the mark for the substitution and the final answer will be forfeited.

5.5 Negative marking

Normally an incorrect answer cannot be correctly motivated if based on a conceptual mistake. If the candidate is therefore required to motivate in QUESTION 3.2 the answer given to QUESTION 3.1, and QUESTION 3.1 is incorrect, no marks can be awarded for QUESTION 3.2. However, if the answer for, for example, QUESTION 3.1 is based on a calculation, the motivation for the incorrect answer in QUESTION 3.2 should be considered.

6. MARKING GUIDELINES: PAPER 2

6.1 Calculations

6.1.1 Marks will be awarded for: correct formula, correct substitution, correct answer with unit.

6.1.2 No marks will be awarded if an **incorrect or inappropriate formula is used**, even though there may be relevant symbols and applicable substitutions.

6.1.3 When an error is made during **substitution into a correct formula**, a mark will be awarded for the correct formula and for the correct substitutions, but **no further marks** will be given.

6.1.4 If **no formula** is given, but **all substitutions are correct**, the candidate will **forfeit one mark**.

6.1.5 Marks are only awarded for a formula if a **calculation has been attempted**, i.e. substitutions have been made or a numerical answer has been given.

6.1.6 Marks can only be allocated for substitutions when values are substituted into formulae and not when listed before a calculation starts.

6.1.7 The final answer to all calculations, when not specified in the question, must be rounded off to a minimum of TWO decimal places.

6.1.8 If a final answer to a calculation is correct, full marks will not automatically be awarded. Markers will always ensure that the correct/appropriate formula is used and that workings, including substitutions, are correct.

6.1.9 Mathematical manipulations and change of subject of appropriate formulae carry no marks, but if a candidate starts off with the correct formula and then changes the subject of the formula incorrectly, marks will be awarded for the formula and the correct substitutions. The mark for the incorrect numerical answer is forfeited.

Example:

CORRECT	ANSWER (1)	POSSIBLE	ANSWER (2)	POSSIBLE
$n = \dfrac{m}{M}$ ✓ $0,01$✓ $= \dfrac{m}{52}$ ✓ $m = 0,52$ g ✓	$n = \dfrac{m}{M}$ ✓ $0,01$ ✓$= \dfrac{52}{m}$ ✗ $m = 5\ 200$ g ✗	$m = \dfrac{n}{M}$ ✗ $= \dfrac{0,01}{52}$ $= 0,002$ g	$n = \dfrac{m}{M}$ ✓ $m = \dfrac{M}{n}$ ✗ $= \dfrac{52}{0,01}$ ✓ ✓ $= 5\ 200$ g✗	$n = \dfrac{m}{M}$ ✓ $= 0,52$ g ✓
(4)	(2)	(0)	(3)	(2)

6.2 Units

6.2.1 Candidates will only be penalised once for the repeated use of an incorrect unit **within a question**.

6.2.2 Units are only required in the final answer to a calculation.

6.2.3 Marks are only awarded for an answer and not for a unit per se. Candidates will therefore forfeit the mark allocated for the answer in each of the following situations:
- Correct answer + wrong unit
- Wrong answer + correct unit
- Correct answer + no unit

6.2.4 Separate compound units with a multiplication dot, not a full stop, for example $mol \cdot dm^{-3}$. Accept $mol.dm^{-3}$ (or mol/dm^3) for marking purposes.

6.3 General

6.3.1 If one answer or calculation is required, but two are given by the candidate, only the first one will be marked, irrespective of which one is correct. If two answers are required, only the first two will be marked, etc.
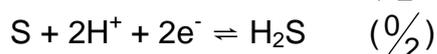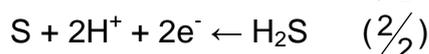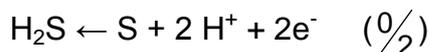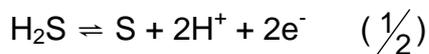
6.3.2 When a chemical **FORMULA** is asked, and the **NAME** is given as answer, the candidate forfeits the marks. The same rule applies when the **NAME** is asked and the **FORMULA** is given.

6.3.3 When redox half-reactions are to be written, the correct arrow should be used.
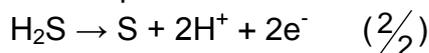
If the equation
$H_2S \rightarrow S + 2H^+ + 2e^-$ $(\frac{2}{2})$
is the correct answer, the marks must be given as follows:

$H_2S \rightleftharpoons S + 2H^+ + 2e^-$ $(\frac{1}{2})$

$H_2S \leftarrow S + 2H^+ + 2e^-$ $(\frac{0}{2})$

$S + 2H^+ + 2e^- \leftarrow H_2S$ $(\frac{2}{2})$

$S + 2H^+ + 2e^- \rightleftharpoons H_2S$ $(\frac{0}{2})$

6.3.4 When candidates are required to give an explanation involving the relative strength of oxidising and reducing agents, the following is not accepted:
- Stating the position of a substance on Table 4 only (e.g. Cu is above Mg).
- Using relative reactivity only (e.g. Mg is more reactive than Cu).
- The correct answer would be for instance: Mg is a stronger reducing agent than Cu, and therefore Mg will be able to reduce Cu^{2+} ions to Cu. The answer can also be given in terms of the relative strength as electron acceptors and donors.

6.3.5 One mark is forfeited when the charge of an ion is omitted per equation (not for the charge on an electron).

6.3.6 The error-carrying principle does not apply to chemical equations or half-reactions. For example, if a learner writes the wrong oxidation/reduction half-reaction in the subquestion and carries the answer to another subquestion (balancing of equations or calculation of E^θ_{cell}), then the learner will not be credited for this substitution.

6.3.7 In the structural formula of an organic molecule all hydrogen atoms must be shown. Marks will be deducted if hydrogen atoms are omitted.
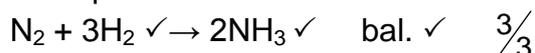
6.3.8 When a structural formula is required, marks will be deducted if the learner writes the condensed formula.

6.3.9 When a IUPAC name is asked and the candidate omits the hyphen(s) (e.g. instead of pent-1-ene or 1-pentene the candidate writes pent 1 ene or 1 pentene), marks will be forfeited.

6.3.10 When a chemical reaction is asked, marks are awarded for correct reactants, correct products and correct balancing.
If only a reactant(s) followed by an arrow, or only a product(s) preceded by an arrow, is/are written, marks may be awarded for the reactant(s) or product(s).
If only a reactant(s) or only a product(s) is/are written, without an arrow, no marks are awarded for the reactant(s) or product(s).

Examples:
$N_2 + 3H_2 \checkmark \rightarrow 2NH_3 \checkmark$ bal. $\checkmark$ $\frac{3}{3}$

$N_2 + H_2 \rightarrow \checkmark$ $\frac{1}{3}$

$\rightarrow NH_3 \checkmark$ $\frac{1}{3}$

$N_2 + H_2$ $\frac{0}{3}$

NH_3 $\frac{0}{3}$

6.4   Positive marking

Positive marking regarding calculations will be followed in the following cases:

6.4.1 Subquestion to subquestion: When a certain variable is calculated in one subquestion (e.g. QUESTION 3.1) and needs to be substituted in another (QUESTION 3.2 or QUESTION 3.3), e.g. if the answer for QUESTION 3.1 is incorrect and is substituted correctly in QUESTION 3.2 or QUESTION 3.3, **full marks** are to be awarded for the subsequent subquestions.

6.4.2 A multistep question in a subquestion: If the candidate has to calculate, for example, current in the first step and gets it wrong due to a substitution error, the mark for the substitution and the final answer will be forfeited.

6.5   Negative marking

Normally an incorrect answer cannot be correctly motivated if based on a conceptual mistake. If the candidate is therefore required to motivate in QUESTION 3.2 the   answer given to QUESTION 3.1, and QUESTION 3.1 is incorrect, no marks can be awarded for QUESTION 3.2. However, if the answer for, for example, QUESTION 3.1 is based on a calculation, the motivation for the incorrect answer in QUESTION 3.2 could be considered.

7.   CONCLUSION

This Examination Guidelines document is meant to articulate the assessment aspirations espoused in the CAPS document. It is therefore not a substitute for the CAPS document which educators should teach to.

Qualitative curriculum coverage as enunciated in the CAPS cannot be over-emphasised.