Periodic Table:

1. Describe the Periodic Table as an arrangement of the elements in order of increasing proton number
2. Describe how the position of an element in the Periodic Table is related to proton number and electronic structure
3. Describe the relationship between group number and the ionic charge of an element

Periodic Table:

- Method of arranging and classifying the elements
- Arranged in order of proton number (or atomic number)
- Vertical columns are Groups and horizontal rows are Periods
- From left to right in one period:
    - Elements go from metals to non-metals
    - Usually the size of the atom becomes smaller (smaller atomic radius)
- Atoms of elements in the same period have the same number of shells
- Elements of the same group
    - Have same number of electrons in their outer shell
    - Similar chemical properties and bonding
    - Form ions and compounds with similar formula (similar charges on ions)

4. Understand the relationship between number of protons and shielding electrons
5. Explain trends in ionization energies across a period of the Periodic Table
6. Describe and explain qualitatively the variation in atomic radius of the elements

Trends in the Periodic Table:

- Nuclear Charge
    - Sum of all the protons in the nucleus of the atom (same as the proton number)
    - Nuclear charge and shielding electrons will constitute the nuclear attraction (which becomes stronger across a period, since the nuclear charge increases while shielding electrons remain the same)
- Shielding Electrons
    - Electrons not in the valence shell of the atom, shielding the outer shell electrons from the attraction force of the nucleus charge
    - Number of shielding electrons remains the same across a Period, increases down a Group
- Atomic Radii
    - Atomic radii increases down a group and generally decreases across a period
    - Down the group, atoms get larger since an additional shell of electrons will be added
    - Across a Period, there is an increase in positive charges while shielding electrons remain constant, so the valence electrons are increasingly pulled towards the nucleus, decreasing the atomic radii
- Ionization Energies
    - Measure of the energy required to remove consecutive electrons from one mole of the gaseous atoms of an element
    - Decreases down a Group: more electron shells means the valence electrons are further away from the nucleus, so the force of attraction between the valence electron and nucleus decreases, so the ionization energy decreases
    - Across Period: nuclear charge increases but shielding electrons remain constant, so the force of attraction between nucleus and valence electrons increase, increasing ionization energy
- Melting Points
    - Factor: size of the force of attraction between the particles in the substance

- Metals:
    - Attraction between positive nuclei and delocalized free electrons
    - Decreases down the group, since metals with larger atoms have delocalized electrons further away from positive nuclei, so the force of attraction becomes weaker, so there is a lower melting point (to break the bonds)
- Non-metals:
    - Force of attraction between molecules
    - Increases down the group, since molecules become larger and there is a greater area for interaction and intermolecular forces of attraction increase, thus more energy is needed to break the bonds, higher melting point

7. <u>Describe the change from metallic to non-metallic character, as well as the structure of oxides formed (ionic or molecular structure) form left to right across a period of the Periodic Table</u>
8. <u>Describe the relationship between group number, number of valence electrons and metallic character</u>
9. <u>State that metals form oxides that are basic or amphoteric while non-metals form oxides that are acidic or neutral</u>

Metals and Non-Metals:

- Metals are usually solids at rtp, with high melting and boiling point (non-metals are often gases)
- Metals are good conductors of heat and electricity (non-metals are poor conductors, except graphite)
- Metals are often shiny, ductile, malleable and have great tensile strength (non-metals dull and soft)
- Metal oxides are basic or amphoteric (non-metal oxides are acidic or neutral)
- Metals form positive ions (non-metals usually form negative ions)

10. *<u>Predict the properties of elements in Group I, VII or 0 and the transition elements using the Periodic Table</u>*
11. *<u>Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity</u>*
12. *<u>Deduce the nature, possible position in the Periodic Table, and identity of unknown elements from given information of physical and chemical properties</u>*

13. <u>Describe lithium, sodium and potassium in Group I as a collection of relatively soft, low density metals showing a trend in melting point and their reaction with water</u>

Group I Elements:

Element	Symbol	Proton Number	Electronic Structure	Melting point/Celsius	Density/ gcm^-3	Physical properties
Lithium	Li	3	2.1	191	0.534	Soft and silvery
Sodium	Na	11	2.8.1	98	0.971	
Potassium	K	19	2.8.8.1	64	0.862	
Rubidium	Rb	37	2.8.18.8.1	39	1.532	
Caesium	Cs	55	2.8.18.18.8.1	29	1.873	

- A group of highly reactive metals known as alkali metals
- The density increases and the melting point decreases down Group I
- Sodium:
    - Upon reaction with water, sodium melts into a ball and darts around on the surface of water
    - A flame can be seen
    - $2Na + H_2O \rightarrow 2NaOH + H_2$

- The hydrogen gas is observed as effervescence, while NaOH is soluble and dissociates into ions, thus explaining why the solution is alkali
- Test for hydrogen gas: 'pop' sound upon inserting lighted splint

- Upon reaction of alkali metals with Universal Indicator, the solution changes from green to blue/purple in colour, and this indicates that the solution is alkali
  - That is the reason why Group I metals are alkali metals
  - Formed alkali compounds (concentration of OH- is higher than that of H+)
- The reactivity of the elements increases down the group
  - Elements lose electrons more readily and larger atomic radius means that electrons can be lost more easily from the valence shell (thus the reactions become more vigorous)

14. <u>Describe chlorine, bromine and iodine in Group VII as a collection of diatomic non-metals showing a trend in colour, state their displacement reactions with solutions of other halide ions</u>

Group VII Elements:

Element	Symbol	Atomic Number	Electronic Structure	Physical state at rtp	Chemical Formula	Colour of Substance
Fluorine	F	9	2.7	Gas	F_2	Pale yellow
Chlorine	Cl	17	2.8.7	Gas	Cl_2	Greenish-yellow
Bromine	Br	35	2.8.18.7	Liquid	Br_2	Reddish-brown
Iodine	I	53	2.8.18.18.7	Solid	I_2	Purple
Astatine	At	85	2.8.18.32.18.7	Solid	At_2	Black

- The colour of substance becomes darker and the state changes from gas to liquid to solid down the group
- Halogens are more soluble in organic solvents than water, and dissolving in organic solvent produces a distinctive coloured solution
  - Iodine is brown in aqueous solution, but violet in organic solvent
    - Iodine is only sparingly soluble in water
    - Blue litmus paper will remain blue
    - Reason: non-polar molecule with high Mr: improved stability of molecule (id-id weak intermolecular bonds)
  - Chlorine water is greenish yellow, and turns blue litmus paper red before it bleaches
    - Reason: weak acid is formed first, then becomes hypochlorous acid (colourless)
  - Bromine water is brown, and turns blue litmus paper red and then bleaches more slowly
- The reactivity of the halogens decreases down the group
  - Fluorine reacts with aluminium and the mixture explodes with blinding white flash
  - Chlorine reacts with aluminium and the heated aluminium sparked and burnt upon coming into contact with chlorine
- A more reactive halogen will displace a less reactive halogen from its salt solution
  - $Cl_2 + 2NaI \rightarrow 2NaCl + I_2$
  - Thus the strongest oxidizing agent is fluorine while the strongest reducing agent is astatine (since F takes electrons away from all other ions, thus oxidizing these ions)

15. <u>Describe the elements in Group 0 as a collection of monatomic elements that are chemically unreactive and hence important in providing an inert atmosphere</u>
    a. <u>Argon and neon in light bulbs</u>
    b. <u>Helium in balloons</u>
    c. <u>Argon in the manufacture of steel</u>
16. <u>Describe the lack of reactivity of the noble gases in terms of their electronic structures</u>

Group 0 Elements:

- A collection of gases known as noble gases
- Monoatomic gases (made up of single atoms) and are generally unreactive
- The melting point increases down the group
  - He has the lowest melting point
- The density increases down the group
  - Rn is the densest noble gas
- Argon is used to fill electric light bulbs and provide an inert atmosphere in steel production
- Helium is used to fill weather balloons (low density allows balloons to rise above atmosphere)

17. <u>Describe the central block of elements (transition metals) as metals having high melting points, high density, variable oxidation states in their compounds and forming coloured compounds</u>
18. <u>State the use of transition metals or their compounds as catalysts</u>

Transition Metals:

Element	Symbol	Atomic Number	Melting Point	Density	Oxidation States	Reaction
Manganese	Mn	25	1245	7.4	+7, +6, +4, +3, +2, 0	MnO2 catalyses decomposition of H2O2
Iron	Fe	26	1536	7.9	+3, +2, 0	Fe catalyses synthesis of ammonia
Nickel	Ni	28	1453	8.9	+3, +2, 0	Ni catalyses hydrogenation
Copper	Cu	29	1023	8.9	+2, +1, 0	CuSO4 catalyses reaction of acid and zinc
Osmium	Os	76	3054	22.6	+4, +2, +1, 0, -2	NIL
Platinum	Pt	78	1772	21.5	+6, +5, +4, +2, 0	Pt catalyses cracking of petroleum

- High melting and boiling points (metallic bonding)
- High density (denser than Group I metals)
- Harder than Group I metals (alkali metals can be cut with razor blade)
- Variable oxidation states in compounds (easily switch its colour and oxidation state)
- Form coloured compounds, basic and amphoteric oxides
- Usually used as catalysts in biological and industrial reactions (large surface area with active sites)
  - Can also act like enzymes and indicators
- Explanation: outermost orbital is d orbital (with up to 10 electrons) compared to Group I (1 electron)

19. <u>*Chemical and physical properties of the elements appear in repeating patterns organized by the Periodic Table*</u>

Group I and Transition Metals:

- Similarities
  - Shiny and malleable and ductile
  - Good conductors of electricity
  - Due to general properties as metals (metallic bonding)
  - Most react with acids to form salt and hydrogen
  - Most react with non-metals to form ionic compounds
  - Most burn in oxygen to form basic or amphoteric oxides
- Differences:

    - Atomic radius, melting/boiling point, density and hardness
    - Group I elements have only one oxidation state, transition metals have multiple
    - Group I elements form white/colourless compounds, transition metals have various colours
    - Group I compounds are all soluble, while not all transition metal compounds are soluble
    - Group I elements have no catalytic properties, transition metals catalyze many reactions
- Reasons: Group I metals can make use of only s-orbital electrons for bonding, while transition metals use their outermost d-orbital electrons for bonding