Class 11 Classification of Elements and Periodicity in Properties Important Questions with Answers

Short Answer Type Questions

1. Explain why the electron gain enthalpy of fluorine is less negative than that of chlorine.

Answer.
This is due to the small size of the fluorine atom. As a result of the strong interelectronic repulsions in fluorine's relatively small 2p orbitals, the incoming electron does not experience much attraction.

2. All transition elements are d-block elements, but all d-block elements are not transition elements. Explain.

Answer.
All transition metals are d-block elements, but not all d-block elements are transition elements because all d-block elements that do not have completely filled d- orbitals are not counted as transition elements, making such elements exceptional. Zn, Cd, and Hg are a few examples.

3. Identify the group and valency of the element having atomic number 119. Also, predict the outermost electronic configuration and write the general formula of its oxide.

Answer.
For element having atomic number 119:
Group - 1, Valency - 1, Outermost electronic configuration - $8s^1$ and the general formula of the oxide will be M_2O.

4. Ionisation enthalpies of elements of second period are given below :
Ionisation enthalpy/ k cal mol^{-1} : 520, 899, 801, 1086, 1402, 1314, 1681, 2080.
Match the correct enthalpy with the elements and complete the graph given in Fig. 3.1.
Also, write symbols of elements with their atomic number.

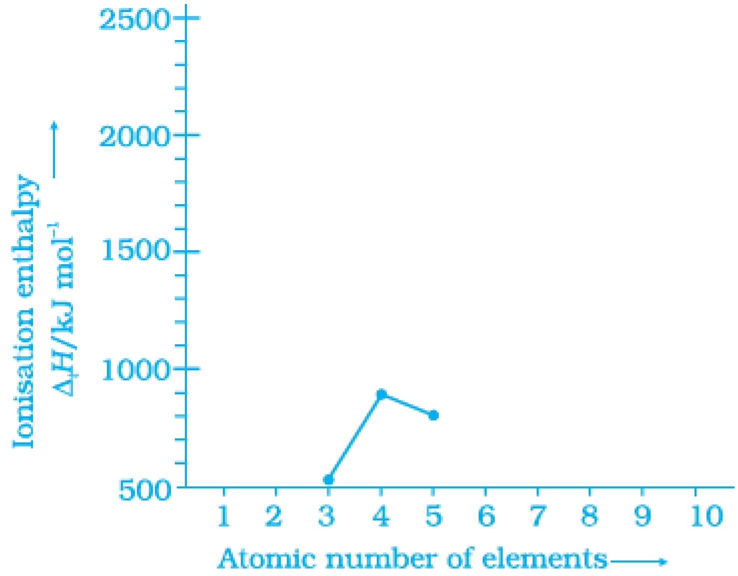
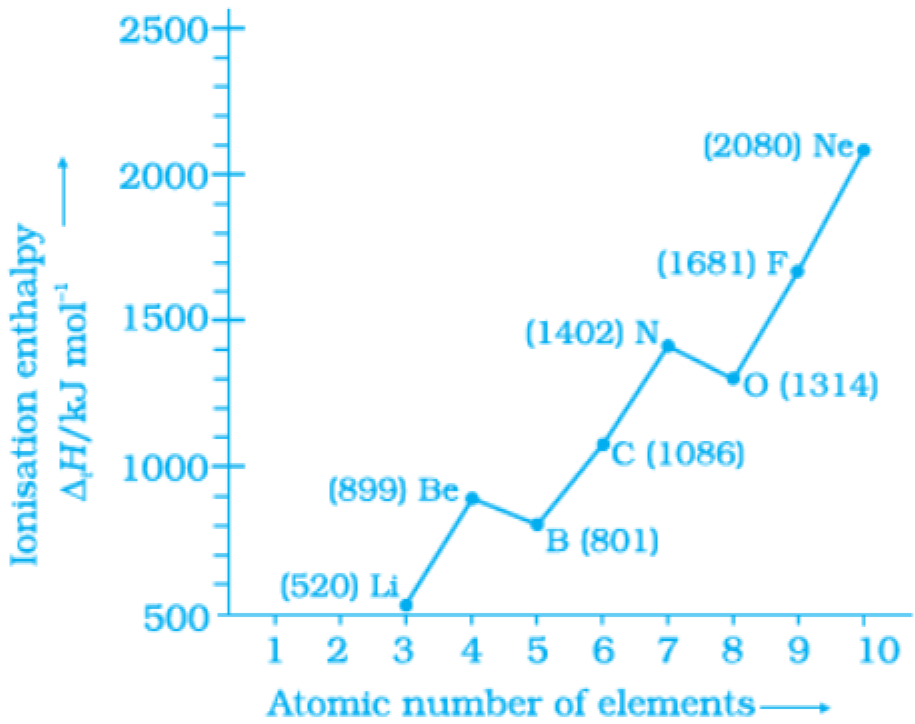


Fig. 3.1

Answer.

The order of ionisation enthalpy of second period elements is: Li < B < Be < C < O < N < F < Ne

The correct enthalpy with the elements and the complete graph is given as follows-



5. Among the elements B, Al, C and Si,
(i) which element has the highest first ionisation enthalpy?
(ii) which element has the most metallic character?
Justify your answer in each case.

Answer.
Among the elements, B, Al, C and Si
(i) The element that has the highest first ionisation enthalpy is C.
(ii) The element that has the most metallic character is Al.

6. Write four characteristic properties of p-block elements.

Answer.
The four most important properties of p-block elements are as follows:
(a) Both metals and nonmetals are present in p-block elements, but the number of nonmetals is much greater than that of metals. Furthermore, within a group, the metallic character increases from top to bottom, while the non-metallic character increases from left to right along a period in this block.
(b) Their ionisation enthalpies are higher than those of s-block elements.
(c) They mostly combine to form covalent compounds.
(d) Some of them have compounds with multiple (variable) oxidation states. In a period, their oxidising character increases from left to right, while their reducing character increases from top to bottom.

7. Choose the correct order of atomic radii of fluorine and neon (in pm) out of the options given below and justify your answer.
(i) 72, 160
(ii) 160, 160
(iii) 72, 72
(iv) 160, 72

Answer.
The atomic radius of F is usually expressed in terms of covalent radius, whereas the atomic radius of neon is usually expressed in terms of van der Waals radius. An element's van der Waals radius is always greater than its covalent radius. As a result, the atomic radius of F is less than the atomic radius of Ne (F = 72 pm, Ne = 160 pm).

8. Illustrate by taking examples of transition elements and non-transition elements that oxidation states of elements are largely based on electronic configuration.

Answer.
An element's oxidation state is determined by its electronic configuration.
S-block elements:

- Group 1 elements- General electronic configuration of the valence shell ns^1. Oxidation states =+1.
- Group 2 elements- General electronic configuration of the valence shell is ns^2. Oxidation state =+2

P-block elements.

- Group 13elements: General electronic configuration of a valence shell is ns^2, np^1. The oxidation states are +3 and +1.
- Group 14 elements: General electronic configuration of the valence shell is ns^2, np^2. Oxidation states are +4 and +2.
- Group 15 elements: General valence shell electronic configuration ns^2, np^3. Oxidation states are -3,+3, and +5. Nitrogen exhibits +1,+2,+4 oxidation states as well.
- Group16 elements General electronic configuration of the valence shell ns^2, np^4. Oxidation states are -2,+2,+4 and +6.
- Group 17: General electronic configuration of the valence shell is ns^2, np^5. Oxidation state is -1. Cl, Br, and I also have +1,+3,+5, and +7 oxidation states.
- Group 18: General electronic configuration of the valence shell is ns^2np^6 . In bonding, the oxidation state = zero.
- A transition metal's various oxidation states are caused by the presence of (n-1)d and outer ns electrons. $(n-1)d^{1-10}$, ns^{1-2} is a general electronic configuration. The most common oxidation states for these elements are +2 and +3.

9. Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.

Answer.

The outermost electronic configuration of nitrogen is $2s^2$ $2p_x^1$, $2p_y^1$, $2p_z^1$ whereas that of oxygen is $2s^2$ $2p_x^2$, $2p_y^1$, $2p_z^1$.

Since oxygen acquires a stable configuration, i.e., $2p^3$, by removing one electron from the 2p-orbital, it has a lower ionisation enthalpy than nitrogen. In the case of nitrogen, however, due to its stable configuration, it is difficult to remove one of the three 2p-electrons.

10. First member of each group of representative elements (i.e., s and p-block elements) shows anomalous behaviour. Illustrate with two examples.

Answer.

The first member of each group of representative elements (i.e., the s- and p- block elements) exhibits anomalous behaviour due to:

(i) small size

(ii) higher ionisation enthalpy

(iii) higher electronegativity

(iv) the absence of d- orbitals.

For example, in s - block elements, lithium behaves differently than the other alkali metals.
(a) Lithium compounds have a high covalent character. Alkali metal compounds are predominantly ionic.
(b) Lithium nitride is formed when lithium reacts with nitrogen, whereas other alkali metals do not form nitrides.
In p- block elements, the first member of each group has four orbitals in their valence shell, one 2s orbital and three 2 p orbitals. As a result, these elements have a maximum covalency of four, whereas other members of the same or different group have a maximum covalency that is greater than four due to the availability of vacant d - orbitals.

11. p-Block elements form acidic, basic and amphoteric oxides. Explain each property by giving two examples and also write the reactions of these oxides with water.

Answer.
Due to their various properties, p - block elements produce acidic, basic, and amphoteric oxides:

- The higher an element's electronegativity, the more acidic its oxide.
  For example- Boron has an electronegativity of -2, carbon has an electronegativity of 2.5, and nitrogen has an electronegativity of 3. As a result, the order of acidic character of B, C, and N oxides is $B_2O_3 < CO_2 < N_2O_3$
- If the ionisation enthalpy of an element is high, it will form acidic oxide; if it is intermediate, it will form amphoteric oxide; and if it is low, it will form basic oxides.
  For example, the ionisation enthalpy of boron is 800 while that of carbon is 1086.5, implying that carbon oxide is more acidic than boron oxide.
- The oxides of the first element in each group in the p - block are more acidic than the oxides of other elements. As we move down the group, the acidic character decreases, followed by elements that form amphoteric oxides and then basic oxides.
  For example- In the Boron family, B forms a weak acidic oxide, while Al, Ga, and In form amphoteric oxides, and Tl forms a strong basic oxide.

Reactions of some of the oxides with water:

- Acidic Oxides:
  $B_2O_3 + 3H_2O \rightarrow 2H_3BO_3$
- Basic Oxides:
  $Tl_2O + H_2O \rightarrow 2TlOH$
- Amphoteric Oxides are insoluble in water and thus reacts with acid and base:
  $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$
  $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$

12. How would you explain the fact that first ionisation enthalpy of sodium is lower than that of magnesium but its second ionisation enthalpy is higher than that of magnesium?

Answer.

The electronic configuration of both the atoms are as follows:
Na: $[Ne]3s^1$
Mg: $[Ne]3s^2$
The ion formed after removing one electron from the sodium atom takes on the configuration of an inert gas, neon, whereas Mg retains one electron. As a result, the first ionisation energy of Na is lower than that of Mg. High energy is required to remove an electron from a noble gas configuration.
As a result, the second ionisation enthalpy of calcium is greater than that of magnesium.

13. What do you understand by exothermic reaction and endothermic reaction?
Give one example of each type.

Answer.
Exothermic reactions are those in which heat is produced. The formation of ammonia from nitrogen and hydrogen, for example, is an exothermic reaction.
$N_2 + 3H_2 \rightarrow 2NH_3$, ΔH = - 46.11KJ/mol.
Endothermic reactions are ones in which heat is absorbed. The dissociation of ammonia into nitrogen and hydrogen, for example, is an endothermic reaction.
$2NH_3 \rightarrow N_2 + 3H_2$, ΔH = +91.8KJ/mol.

14. Arrange the elements N, P, O and S in the order of-
(i) increasing first ionisation enthalpy.
(i) increasing non-metallic character.
Give reason for the arrangement assigned.

Answer.
(i) Due to the extra stable exactly half-filled 2p-orbitals, the ionisation enthalpy of nitrogen ($1s^2,2s^2,2p^3$) is greater than that of oxygen ($1s^2,2s.2,2p^4$). Likewise, the ionisation enthalpy of phosphorous ($1s^2,2s^2,2p^6,3s^2,3p^3$) is greater than that of sulphur ($s^2,2s^2,2p^6,3s^2,3p^3$)
Ionisation enthalpy decreases with decreasing atomic size as one moves down the group.
As a result, the increasing order of first ionisation enthalpy is S < P < O < N
(ii) Nonmetallic character increases across a period (left to right), but decreases as one moves down the group.
As a result, the increasing non-metallic order is P < S < N < O.

15. Explain the deviation in ionisation enthalpy of some elements from the general trend by using Fig. 3.2.

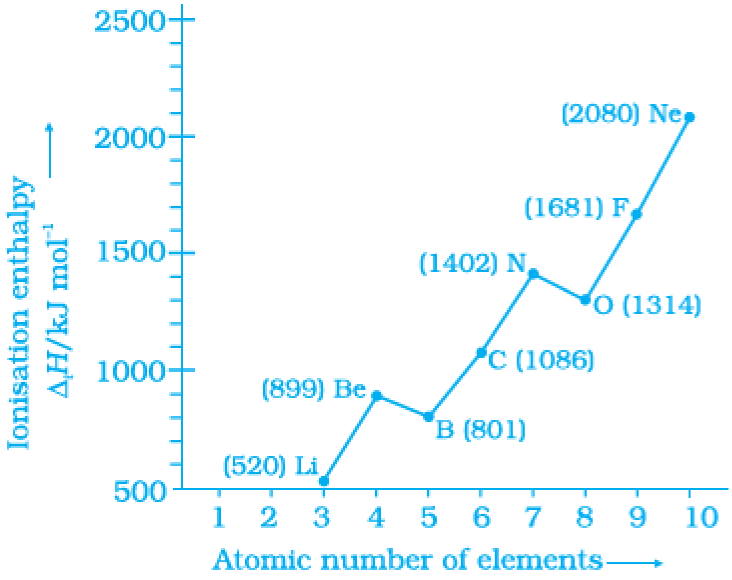


Fig. 3.2

Answer.
The ionisation enthalpy of some elements deviates from the general trend, as shown in the figure. The first ionisation enthalpy of B is lower than that of Be, whereas the first ionisation enthalpy of nitrogen is higher than that of O.

16. Explain the following:

(a) Electronegativity of elements increase on moving from left to right in the periodic table.
(b) Ionisation enthalpy decrease in a group from top to bottom?

Answer.

(a) Due to the general decrease in atom size and increase in nuclear charge, the electronegative of elements increases as one moves from left to right in the periodic table.
(b) The ionisation enthalpy of a group decreases from top to bottom due to the increase in atomic size caused by the addition of a new shell.

17. How does the metallic and non metallic character vary on moving from left to right in a period?

Answer.
Moving from left to right in a period, metallic character decreases and nonmetallic character increases. This is due to the increase in effective nuclear charge which causes an increase in ionisation enthalpy and electron gain enthalpy.

18. The radius of Na^+ cation is less than that of Na atom. Give reason.

Answer.
Since Na+ is formed by losing one energy shell, its radius is smaller than that of Na atom.
Na - $1s^2, 2s^2, 2p^6, 3s^1$
Na^+ - $1s^2, 2s^2, 2p^6$

19. Among alkali metals which element do you expect to be least electronegative and why?

Answer.
Due to the general increase in size, electronegativity decreases from top to bottom in a group. As a result, caesium is the least electronegative element.

Matching Type Questions

1. Match the correct atomic radius with the element.

Element	Atomic radius (pm)
Be	74
C	88
O	111
B	77
N	66

Answer.

Element	Atomic radius (pm)
Be	111
C	66
O	77
B	88
N	74

2. Match the correct ionisation enthalpies and electron gain enthalpies of the following elements.

Elements		ΔH_1	ΔH_2	$\Delta_{eg} H$
(i) Most reactive non-metal	A.	419	3051	- 48
(ii) Most reactive metal	B.	1681	3374	- 328
(iii) Least reactive element	C.	738	1451	- 40
(iv)Metal forming binary halide	D.	2372	5251	+ 48

Answer.

Elements		ΔH_1	ΔH_2	$\Delta_{eg} H$
(i) Most reactive non-metal	B.	1681	3374	- 328
(ii) Most reactive metal	A.	419	3051	- 48
(iii) Least reactive element	D.	2372	5251	- 48
(iv)Metal forming binary halide	C.	738	1451	+ 40

3. Electronic configuration of some elements is given in Column I and their electron gain enthalpies are given in Column II, Match the electronic configuration with electron gain enthalpy.

Column I Electronic configuration	Column II Electron gain enthalpy/KJ mol^{-1}
(i) $1s^2\ 2s^2\ 2p^6$	(A) - 53
(ii) $1s^2\ 2s^2\ 2p^6\ 3s^1$	(B) - 328
(iii) $1s^2\ 2s^2\ 2p^5$	(C) - 141
(iv) $1s^2\ 2s^2\ 2p^4$	(D) +48

Answer.

Column I Electronic configuration	Column II Electron gain enthalpy/KJ mol^{-1}
(i) $1s^2\ 2s^2\ 2p^6$	(D) - 328
(ii) $1s^2\ 2s^2\ 2p^6\ 3s^1$	(A) - 53
(iii) $1s^2\ 2s^2\ 2p^5$	(B) +48
(iv) $1s^2\ 2s^2\ 2p^4$	(C) - 141

Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of reason (R) is given. Choose the correct option out of the choices given below each question.

1. Assertion (A): Generally, ionisation enthalpy increases from left to right in a period.
Reason (R): When successive electrons are added to the orbitals in the same principal quantum level, the shielding effect of the inner core of electrons does not increase very much to compensate for the increased attraction of the electron to the nucleus.

(i) Assertion is correct statement and reason is wrong statement.
(ii) Assertion and reason both are correct statements and reason is correct explanation of assertion.
(iii) Assertion and reason both are wrong statements.
(iv) Assertion is wrong statement and reason is correct statement.

Answer.
The correct answer is (ii) Assertion and reason both are correct statements and reason is correct explanation of assertion.
Ionisation enthalpy increases over time as effective nuclear charge increases and atomic size decreases.

2. Assertion (A): Boron has a smaller first ionisation enthalpy than beryllium.
Reason (R): The penetration of a 2s electron to the nucleus is more than the 2p electron hence 2p electron is more shielded by the inner core of electrons than the 2s electrons.

(i) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
(ii) Assertion is correct statement but reason is wrong statement.
(iii) Assertion and reason both are correct statements and reason is correct explanation for assertion.
(iv) Assertion and reason both are wrong statements.

Answer.
The correct answer is (iii) Assertion and reason both are correct statements and reason is correct explanation for assertion.
The electron removed from the beryllium atom during ionisation is from the s-orbital, whereas the electron removed from the boron atom is from the p-orbital, and the penetration of the 2s electron to the nucleus is greater than that of the 2p electron, so the 2p electron of boron is more shielded from the nucleus than the 2s electron.

3. Assertion (A): Electron gain enthalpy becomes less negative as we go down a group.
Reason (R): Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

(i) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
(ii) Assertion and reason both are correct statements and reason is correct explanation for assertion.
(iii) Assertion and reason both are wrong statements.
(iv) Assertion is wrong statement but reason is correct statement.

Answer.
The correct answer is (ii) Assertion and reason both are correct statements and reason is correct explanation for assertion.
As an atom's size increases down the group, its electron gain enthalpy decreases. This is due to the fact that the screening effect within a group increases as one moves downward, and the added electron would be further away from the nucleus.

Long Answer Type Questions

1. Discuss the factors affecting electron gain enthalpy and the trend in its variation in the periodic table.

Answer.

Factors influencing electron gain enthalpy include-

(i) Nuclear charge: As the nuclear charge increases, the electron gain enthalpy becomes more negative. If the nuclear charge is high, there is a greater attraction for the incoming electron.
(ii) Atomic size: As the atom's size increases, so does the distance between the nucleus and the incoming electron, resulting in less attraction. As a result, as the size of the element's atom increases, the electron gain enthalpy becomes less negative.
(iii) Electronic configuration: Elements with stable electronic configurations of half-filled and completely filled valence subshells have a very low tendency to accept additional electrons, resulting in less negative electron gain enthalpies.

Variations in electron gain enthalpies in the periodic table

In general, electron gain enthalpy becomes more negative from left to right in a period and less negative from top to bottom in a group.

(a) Downward variation within a group: Moving down a group increases the size and nuclear charge. However, the effect of increasing atomic size is much more pronounced than that of increasing nuclear charge, so the additional electron feels less attraction by the large atom. As a result, the electron gain enthalpy decreases. This is evident from the decrease in electron gain enthalpy when transitioning from chlorine to bromine and then to iodine.
(b) Periodic variation: As one moves across a period, the size of the atom decreases and the nuclear charge increases. Because both of these factors increase the attraction for the incoming electron, electron gain enthalpy becomes more negative in a period from left to right. However, there are some anomalies in the overall trend. These are primarily due to certain atoms' stable electronic configurations.

Important Trends in Electron Gain Enthalpies

The electron gain enthalpies of elements have some important characteristics. They are as follows:
(i) The negative electron gain enthalpies of halogens are the highest.
(ii) Noble gases have positive electron gain enthalpy values, whereas Be, Mg, N, and P have nearly zero.
(iii) Fluorine's electron gain enthalpy is unexpectedly less negative than chlorine's.

2. Define ionisation enthalpy. Discuss the factors affecting ionisation enthalpy of the elements and its trends in the periodic table.

Answer.
Ionization Enthalpy: The amount of energy required to remove an e from an isolated gaseous atom in its gaseous state is defined as an element's ionisation enthalpy.
The following factors influence ionisation enthalpy:

1. Atom size: The larger the atomic size, the lower the value of ionisation enthalpy. The outer e^- are far away from the nucleus in large atoms, so the force of attraction with which they are attracted by the nucleus is less and thus they can be easily removed.
   Ionization enthalpy $\propto$ 1/atomic size
2. Screening Effect: Because the screening effect reduces the force of attraction towards the nucleus, the outer e can be easily removed.
   Ionization enthalpy $\propto$ 1/Screening effect
3. Nuclear charge: The Ionization enthalpy increases as nuclear charge increases among atoms with the same number of energy shells because of the force of attraction towards the nucleus

increases.
Ionization enthalpy ∝ nuclear charge
4. Half-filled and fully-filled orbitals: Because atoms with half-filled and fully-filled orbitals are more stable, it takes more energy to remove an electron from such atoms. In the case of such an atom, the ionisation enthalpy is somewhat higher than expected. Ionization enthalpy ∝ stable electronic configuration
5. Orbital shape: The s-orbital of the same orbit is closer to the nucleus than the p-orbital. As a result, removing an electron from a p-orbital is easier than from an s-orbital. The shape for orbitals: s > p > d > f

Variation of ionisation enthalpy in the periodic table
In general, as atomic size increases, the ionisation energy decreases down the group. The ionisation energy, on the other hand, increases across the period from left to right, because of a decrease in atomic size from left to right.

3. Justify the given statement with suitable examples— “the Properties of the elements are a periodic function of their atomic numbers”.

Answer.
This statement means that when elements are arranged in increasing atomic number order, the similarity of their properties appears at regular intervals.

The electronic configuration of the elements, for example, repeats itself at regular intervals, resulting in elements with the same number of valence electrons and thus similar chemical properties.
- The elements in the first group are strongly metallic and electropositive, forming monovalent ions such as Na^+, Li^+, K^+, and so on. All alkali metals have similar properties due to their similar outermost shell electronic configuration.
- Similarly, all elements in the 17th group (halogens) have the same outermost shell electronic configuration, i.e., $ns^2\ np^5$, and thus have similar properties.

4. Write down the outermost electronic configuration of alkali metals. How will you justify their placement in group 1 of the periodic table?

Answer.
Alkali metals' outermost electronic configuration is ns^1.
All elements of group IA (or I), i.e., alkali metals, have the same outer electronic configuration, ns^1, where n denotes the number of principal shells. These electronic configurations are shown in the table below.

Symbol	Atomic Number	Electronic Configuration

Li	3	$1s^2\ 2s^1$
Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
K	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Rb	37	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$
Cs	55	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^6\ 6s^1$
Fr	87	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^{10}\ 6s^2\ 6p^6\ 7s^1$

As a result of their similarity in electronic configuration and properties, all of these elements are placed in group 1 of the periodic table.

5. Write the drawbacks in Mendeleev's periodic table that led to its modification.

Answer.

1. Hydrogen's position: Hydrogen is assigned to group I. It does, however, resemble elements from Group I (alkali metals) as well as elements from Group VILA (halogens). As a result, the position of hydrogen in the periodic table is incorrect.
2. Anomalous pairs: The increasing order of atomic masses was not followed in certain pairs of elements. Mendeleev arranged the elements in these cases based on similarities in their properties rather than the increasing order of their atomic masses. Argon (Ar, atomic mass 39.9), for example, is placed before potassium (K, atomic mass 39.1). Likewise, cobalt (Co, atomic mass 58.9) comes before nickel (Ni, atomic mass 58.6), and tellurium (Te, atomic mass 127.6) comes before iodine (I, atomic mass 126.9). These positions were not justified.
3. Isotopes are atoms of the same element that have different atomic masses but the same atomic number. As a result, according to Mendeleev's classification, these should be classified differently based on their atomic masses. For example, hydrogen isotopes with atomic masses 1, 2, and 3 should be placed in three different locations. Isotopes, on the other hand, do not have their own spot in the periodic table.
4. Several gaps in the periodic table were left because he believed that several elements were yet to be discovered, for example, gallium was not discovered at the time.
5. Position of lanthanoids (or lanthanides) and actinoids (or actinides): The fourteen elements that follow lanthanum (known as lanthanoids, atomic numbers 58-71) and the fourteen elements that follow actinium (known as actinoids, atomic numbers 58-71) are not included separately.

6. In what manner is the long form of periodic table better than Mendeleev's periodic table?
Explain with examples.

Answer.
Due to the following reasons, the long-form periodic table is considered more letter than the Mendeleev's table:

1. All elements in the long-form periodic table are arranged in increasing order of atomic numbers, whereas the table is arranged in increasing order of atomic masses.
2. The position of hydrogen in the long-form periodic table has been justified, whereas there is no such justification in Mendeleev's periodic table.
3. The long-form periodic table considers the filling of electrons in s,p,d, and subshells, whereas the table considers the atomic numbers of the elements.
4. The periodic table is divided into four blocks: s, p, d, and f, whereas the periodic table, has no blocks.
5. Long-form periodic table groups are not further subdivided into subgroups, whereas each group in Mendeleev's periodic table has subgroups A and B.
6. Long-form periodic tables are simple and easy to reproduce, whereas Mendeleev's periodic tables are quite difficult to reproduce.

7. Discuss and compare the trend in ionisation enthalpy of the elements of group 1 with those of group 17 elements.

Answer.

1. As the atom grows in size, the distance between the nucleus and the valence electron grows, resulting in a decrease in the force of attraction between them. As a result, the amount of energy required to remove an electron is reduced.
2. Also, as we move down the group, the shielding effect increases, and thus the force of attraction between the nucleus and the valence electrons decreases even more.
3. As we move down a group, the combined effect of the two results in a decrease in ionisation energy.
4. As a result, the group ionisation energy decreases as we move down in both group 1 and group 17.

The ionisation enthalpy values for elements in groups 1 and 17 are given below.

Group I	First Ionization enthalpies (KJ/mol)	Group 17	First Ionization enthalpies (KJ/mol)
H	1312	F	1681
Li	520	Cl	1255
Na	496	Br	1142
K	419	I	1009

Rb	403	At	917
Cs	374		