Chapter 3: Classification of Elements and Periodicity in Properties

Exercises — Exercises

Q1 What is the basic theme of organisation in the periodic table?

SOLUTION

The periodic table organises elements so that elements with similar physical and chemical properties appear together in the same vertical column (group).

The basic theme is the periodicity of properties: when elements are arranged in order of increasing atomic number, similar properties recur at regular intervals.

This allows systematic study and prediction of properties — elements in the same group have the same number of valence electrons, which governs their chemistry.
Horizontal rows (periods) show how properties change progressively across a series of elements as electrons fill a given principal energy level.
Vertical columns (groups) collect elements with the same outer electronic configuration and therefore similar chemical behaviour.

Answer: The basic theme is periodicity: elements with similar outer electronic configurations (and hence similar properties) recur at regular intervals when arranged by increasing atomic number, allowing them to be grouped in vertical columns.

Q2 Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?

SOLUTION

Mendeleev used atomic mass as the basis for classifying elements.
He arranged elements in order of increasing atomic mass and noted that elements with similar properties recurred periodically.

Mendeleev's Periodic Law: 'The properties of elements are a periodic function of their atomic masses.'

No, he did not always stick to this principle.
In a few cases he placed elements out of strict atomic-mass order to ensure that elements with similar properties fell in the same group.
For example, tellurium (Te, atomic mass 127.6) was placed before iodine (I,

atomic mass 126.9) because Te's properties match Group VIA and I's properties match Group VIIA.

Mendeleev gave priority to chemical similarity over strict atomic-mass order in such cases.

Answer: Mendeleev used atomic mass. He did not always stick to it — in cases of conflict he placed elements by chemical similarity rather than strictly increasing atomic mass (e.g., Te before I).

Q3 What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?

SOLUTION

Mendeleev's Periodic Law (1869): 'The physical and chemical properties of elements are a periodic function of their atomic masses.'

Basis of classification: atomic mass (a measurable but empirical quantity).

Problem: some elements had to be placed out of atomic-mass order; isotopes could not be accommodated.

Modern Periodic Law (Moseley, 1913): 'The physical and chemical properties of elements are a periodic function of their atomic numbers.'

Basis of classification: atomic number (Z = number of protons in the nucleus), which is more fundamental.

Key difference: Mendeleev used atomic mass; the Modern Periodic Law uses atomic number.

Using atomic number removes all the anomalies present in Mendeleev's table (Ar/K, Co/Ni, Te/I pairs all fall correctly by Z).

Atomic number determines the electronic configuration, which in turn determines chemical properties — so the modern law has a deeper theoretical basis.

Answer: Mendeleev used atomic mass as the basis; the Modern Periodic Law uses atomic number (Z). Atomic number is more fundamental — it determines electronic configuration and removes all ordering anomalies present in Mendeleev's table.

Q4 On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

SOLUTION

For the sixth period, the principal quantum number n = 6.

Electrons in the sixth period fill the 6s, 4f, 5d, and 6p subshells in order of increasing energy (Aufbau principle).

Subshells filled during Period 6 and their capacities:

6s subshell: l = 0, m_l = 0 only, two spin states => 2 electrons => 2 elements

4f subshell: l = 3, m_l = -3 to +3 (7 orbitals), 2 electrons each => 14 electrons => 14 elements

5d subshell: l = 2, m_l = -2 to +2 (5 orbitals), 2 electrons each => 10 electrons => 10 elements

6p subshell: l = 1, m_l = -1, 0, +1 (3 orbitals), 2 electrons each => 6 electrons => 6 elements

Total elements in Period 6 = 2 + 14 + 10 + 6 = 32

Verification by atomic numbers: Period 6 runs from Cs (Z = 55) to Rn (Z = 86), giving 86 - 55 + 1 = 32 elements.

Answer: Period 6 fills 6s (2), 4f (14), 5d (10), and 6p (6) subshells, giving 2 + 14 + 10 + 6 = 32 elements.

Q5 In terms of period and group where would you locate the element with Z = 114?

SOLUTION

Write the electronic configuration of Z = 114 by filling subshells in Aufbau order:

[Rn] $5f^{14}$ $6d^{10}$ $7s^2$ $7p^2$

Period: The highest principal quantum number is n = 7 => Period 7.

Block: The last subshell to receive electrons is 7p => p-block.

Group: For p-block elements, Group number = 10 + (electrons in s subshell) + (electrons in p subshell)

= 10 + 2 + 2 = 14

=> Group 14

This element (Flerovium, Fl) is in the same group as C, Si, Ge, Sn, and Pb.

Answer: Z = 114 is in Period 7, Group 14 (p-block). Its configuration is [Rn] $5f^{14}$ $6d^{10}$ $7s^2$ $7p^2$ (element Flerovium, Fl).

Q6 Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

SOLUTION

Period 3 means the outermost electrons are in n = 3 shell.
Group 17 means the element has 7 valence electrons (ns^2 np^5 configuration).

Electronic configuration: [Ne] $3s^2$ $3p^5$
Number of electrons = 10 (Ne core) + 2 + 5 = 17

The element is Chlorine (Cl) with atomic number Z = 17.

Answer: Atomic number = 17 (Chlorine, Cl). Configuration: [Ne] $3s^2$ $3p^5$.

Q7 Which element do you think would have been named by (i) Lawrence Berkeley Laboratory (ii) Seaborg's group?

SOLUTION

Elements are often named after the place or person associated with their discovery.

(i) Lawrence Berkeley Laboratory (at the University of California, Berkeley):
The element named after this laboratory is Lawrencium (Lr, Z = 103), named after Ernest O. Lawrence (inventor of the cyclotron).
The city Berkeley gave its name to Berkelium (Bk, Z = 97).
The laboratory itself is associated with Californium (Cf, Z = 98).
Most directly: the element named for Lawrence Berkeley Lab is Lawrencium (Lr).

(ii) Seaborg's group (Glenn T. Seaborg's research group at Berkeley):
Seaborg and his group are credited with the discovery of many transuranium elements.
The element named in his honour is Seaborgium (Sg, Z = 106).

Answer: (i) Lawrence Berkeley Laboratory: Lawrencium (Lr, Z = 103). (ii) Seaborg's group: Seaborgium (Sg, Z = 106).

Q8 Why do elements in the same group have similar physical and chemical properties?

SOLUTION

Elements in the same group have the same number of valence electrons and the same outer electronic configuration (ns^a np^b for main-group elements).

Since chemical properties are primarily determined by the valence electrons:

- The same number of valence electrons means the same valence (oxidation state tendency).
- The same orbital type (s, p, d) means similar bonding behaviour.
- The same electron-pairing pattern leads to similar types of compounds formed.

Physical properties (metallic/non-metallic character, melting point trends) also show family resemblance because they depend on atomic size, ionization enthalpy, and electron affinity, all of which are governed by the outer electronic configuration.

As we go down a group, only n changes (outer electrons enter higher shells), so properties change gradually but the chemical family character is preserved.

Answer: Elements in the same group have the same outer electronic configuration (same number and type of valence electrons), which determines their chemical bonding, oxidation states, and reactivity — hence similar chemical and physical properties.

Q9 What does atomic radius and ionic radius really mean to you?

SOLUTION

Atomic radius:

The atomic radius is a measure of the size of an atom — roughly, the distance from the nucleus to the outermost shell of electrons.

Because electron clouds have no sharp boundary, the radius is defined operationally:

- Covalent radius: half the distance between two identical atoms bonded covalently (e.g., half of Cl–Cl bond length in Cl_2).
- Van der Waals radius: half the distance between two identical non-bonded atoms in contact (e.g., in a solid noble-gas crystal); always larger than the covalent radius.
- Metallic radius: half the internuclear distance between two adjacent atoms in a

metal lattice.

Ionic radius:

The ionic radius is the effective radius of an ion in an ionic crystal — the distance attributed to that ion in a crystal lattice.

It is determined from X-ray diffraction data for ionic compounds.

Cations are smaller than the parent atom (electrons removed, nuclear charge dominates the remaining electrons more).

Anions are larger than the parent atom (extra electrons increase electron–electron repulsion, expanding the electron cloud).

Both are approximations and depend on the environment (coordination number, adjacent ions), but they provide useful comparative measures of atomic/ionic size.

Answer: Atomic radius is a measure of the size of a neutral atom (defined as half the internuclear distance for identical bonded or non-bonded atoms). Ionic radius is the effective radius of an ion in an ionic crystal, determined by X-ray diffraction. Both are approximations since electron clouds have no sharp boundary.

Q10 How do atomic radius vary in a period and in a group? How do you explain the variation?

SOLUTION

Variation across a period (left to right):

Atomic radius generally decreases across a period.

Reason: As we move across a period, nuclear charge (Z) increases by one at each step while electrons are added to the SAME principal shell (n = constant).

The shielding of outer electrons by core electrons remains nearly constant.

Therefore, the effective nuclear charge ($Z_{eff} = Z$ - shielding constant) increases across a period.

Greater Z_{eff} pulls the outer electrons closer to the nucleus, reducing the atomic radius.

(Noble gases are an exception — their van der Waals radii are larger than the covalent radii of the preceding halogens.)

Variation down a group (top to bottom):

Atomic radius generally increases down a group.

Reason: As we move down a group, a new principal quantum shell is added (n increases by 1).

Although Z also increases, the additional core electron shells shield the valence electrons effectively.

The net effect is that valence electrons experience a relatively smaller effective nuclear charge AND occupy a shell that is further from the nucleus.

Therefore, the atomic radius increases progressively down a group.

Answer: Across a period: radius decreases (increasing Z_eff with constant n pulls electrons inward). Down a group: radius increases (addition of new principal shells more than compensates for the increasing Z, as shielding by core electrons increases).

Q11 What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions. (i) F^- (ii) Ar (iii) Mg^{2+} (iv) Rb^+

SOLUTION

Isoelectronic species are atoms or ions that have the same number of electrons (and the same electronic configuration), even though they may have different nuclear charges.

(i) F^- (fluoride ion): Z = 9, charge = -1, electrons = 9 + 1 = 10
Isoelectronic species (10 electrons): Ne, Na^+, Mg^{2+}, Al^{3+}, O^{2-}
Example: Ne (neon atom)

(ii) Ar (argon atom): Z = 18, electrons = 18
Isoelectronic species (18 electrons): Cl^-, K^+, Ca^{2+}, S^{2-}, P^{3-}
Example: Cl^-

(iii) Mg^{2+}: Z = 12, charge = +2, electrons = 12 - 2 = 10
Isoelectronic species (10 electrons): Ne, Na^+, F^-, Al^{3+}, O^{2-}
Example: Na^+

(iv) Rb^+: Z = 37, charge = +1, electrons = 37 - 1 = 36
Isoelectronic species (36 electrons): Kr, Sr^{2+}, Br^-, Se^{2-}
Example: Kr (krypton atom)

Answer: (i) F^- (10 e): Ne. (ii) Ar (18 e): Cl^-. (iii) Mg^{2+} (10 e): Na^+. (iv) Rb^+ (36 e): Kr.

Q12 Consider the following species: N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+} and Al^{3+}. (a) What is common in them? (b) Arrange them in the order of increasing ionic radii.

SOLUTION

(a) What is common:

Counting electrons:

N^{3-}: Z=7, electrons = 7+3 = 10

O^{2-}: Z=8, electrons = 8+2 = 10

F^{-}: Z=9, electrons = 9+1 = 10

Na^{+}: Z=11, electrons = 11-1 = 10

Mg^{2+}:Z=12, electrons = 12-2 = 10

Al^{3+}:Z=13, electrons = 13-3 = 10

All six species have 10 electrons — they are isoelectronic.

All have the neon electronic configuration ($1s^2\ 2s^2\ 2p^6$).

(b) Order of increasing ionic radii:

For isoelectronic species, all have the same number of electrons.

As nuclear charge (Z) increases, the effective nuclear charge on the same 10 electrons increases.

Greater nuclear charge pulls electrons closer => smaller ionic radius.

Z values: $N^{3-}(7) < O^{2-}(8) < F^{-}(9) < Na^{+}(11) < Mg^{2+}(12) < Al^{3+}(13)$

Therefore, ionic radius decreases as Z increases:

$Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$

Increasing order of ionic radii:

$Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$

Answer: (a) All are isoelectronic — each has 10 electrons (neon configuration). (b) Increasing ionic radius: $Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$ (radius decreases as Z increases for isoelectronic species).

Q13 Explain why cations are smaller and anions larger in radii than their parent atoms?

SOLUTION

Cations (formed by loss of electrons):

When an atom loses one or more electrons to form a cation:

- The nuclear charge (Z) remains the same.
- The number of electrons decreases.

- The electron–electron repulsions decrease, so the remaining electrons are held more tightly by the unchanged nucleus.
- In some cases, the outermost shell is entirely removed, leaving the next inner (smaller) shell as the outermost.

Net result: the effective nuclear charge per electron increases => the electron cloud contracts => cation is smaller than the parent atom.

Example: Na (radius ~186 pm) vs Na^+ (radius ~102 pm).

Anions (formed by gain of electrons):

When an atom gains one or more electrons to form an anion:

- The nuclear charge (Z) remains the same.
- The number of electrons increases.
- Greater electron–electron repulsions cause the electron cloud to expand.
- The same nuclear charge now attracts more electrons; the effective nuclear charge per electron decreases.

Net result: the electron cloud expands => anion is larger than the parent atom.

Example: Cl (radius ~99 pm) vs Cl^- (radius ~181 pm).

Answer: Cations are smaller: the same nuclear charge holds fewer electrons more tightly (and often the outermost shell is lost), so the radius shrinks. Anions are larger: extra electrons increase electron–electron repulsion while the nucleus is unchanged, expanding the electron cloud.

Q14 What is the significance of the terms 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy?

SOLUTION

'Isolated gaseous atom':

The atom must be in the gaseous phase so there are no intermolecular forces or lattice energies from neighbours that would affect the energy required to remove or add an electron.

'Isolated' means the atom is far from other atoms — no interactions with solvent, lattice, or neighbouring molecules.

This allows the measurement of a property intrinsic to the atom alone, making values comparable across elements.

'Ground state':

The atom must be in its lowest energy (most stable) electronic configuration.

If the atom were in an excited state, electrons would already have higher energy, and it would be easier to remove them, giving a lower (misleading) ionization enthalpy.

Using the ground state ensures that we measure the maximum stability of the electron being removed or the maximum driving force for electron addition.

Together, these two conditions establish a common, reproducible reference state for comparing ionization enthalpies and electron gain enthalpies across all elements.

Without these qualifications, measured values would not be comparable because they would depend on the physical state, concentration, and excitation level of each atom.

Answer: 'Isolated gaseous atom' removes intermolecular/lattice effects so the measurement reflects only the atom's intrinsic property. 'Ground state' ensures the electron is in its most stable configuration, giving reproducible, comparable values across elements.

Q15 Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1}.

SOLUTION

Given: Energy of electron in ground state of H atom = -2.18×10^{-18} J

Avogadro's number $N_A = 6.022 \times 10^{23}$ mol^{-1}

The energy of the electron in ground state (n=1) of hydrogen is -2.18×10^{-18} J.

To ionize the atom, the electron must be removed to n = infinity (zero energy).

Energy required per atom = 0 - (-2.18×10^{-18} J) = $+2.18 \times 10^{-18}$ J

For one mole of hydrogen atoms:

$\Delta H_i = 2.18 \times 10^{-18}$ J $\times 6.022 \times 10^{23}$ mol^{-1}

= 2.18 x 6.022 x 10^(23-18) J mol^{-1}

= 13.128×10^5 J mol^{-1}

= 1.313×10^6 J mol^{-1}

= 1312.8 kJ mol^{-1}

Verification: Accepted value of IE for H = 1312 kJ/mol — consistent.

Answer: Ionization enthalpy of atomic hydrogen = 1.313×10^6 J mol^{-1} (approximately 1312 kJ mol^{-1}).

Q16 Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne. Explain why (i) Be has higher ΔiH than B (ii) O has lower ΔiH than N and F?

SOLUTION

(i) Be has higher ionization enthalpy than B:

Be: [He] $2s^2$ — both outer electrons are in the 2s orbital (fully filled subshell).

B: [He] $2s^2\ 2p^1$ — the outer electron to be removed is in the 2p orbital.

The 2p orbital is higher in energy than 2s ($l = 1 > l = 0$ for same n).

Also, the 2p electron in B is effectively shielded by the $2s^2$ electrons.

Therefore, the 2p electron in B is easier to remove than a 2s electron in Be.

Additionally, a fully filled $2s^2$ subshell in Be has extra stability.

Net result: Be has higher ionization enthalpy than B.

(ii) O has lower ionization enthalpy than N and F:

N: [He] $2s^2\ 2p^3$ — each of the three 2p orbitals has one electron (half-filled).

A half-filled subshell ($2p^3$) has extra stability due to exchange energy maximization.

O: [He] $2s^2\ 2p^4$ — one 2p orbital has two paired electrons.

The paired electrons in O experience extra electron–electron repulsion.

This repulsion makes it easier to remove one electron from O than from N.

Therefore, O has lower ionization enthalpy than N.

Compared to F: F has a higher nuclear charge (Z=9 vs Z=8 for O).

Despite its $2p^5$ configuration, F's higher Z means its electrons are held more tightly.

Hence F has higher ionization enthalpy than O.

Net result: N > O (half-filled stability in N) and F > O (higher Z in F).

Answer: (i) Be ($2s^2$, fully filled subshell) has higher ionization enthalpy than B ($2p^1$) because a fully filled 2s subshell is more stable and 2p electrons are higher in energy and more shielded. (ii) O ($2p^4$) has lower ionization enthalpy than N ($2p^3$ half-filled, extra stability) and F (higher Z, electrons held more tightly).

Q17 How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?

SOLUTION

Electronic configurations:

Na: [Ne] $3s^1$ — one valence electron

Mg: [Ne] $3s^2$ — two valence electrons

First ionization enthalpy (IE_1):

IE_1 of Na (496 kJ/mol) < IE_1 of Mg (738 kJ/mol)

Mg has higher nuclear charge ($Z=12$) than Na ($Z=11$).

Both have their outermost electron in the 3s subshell.

Mg's 3s electrons experience greater effective nuclear charge => harder to remove.

Hence Mg has higher IE_1 than Na.

Second ionization enthalpy (IE_2):

IE_2 of Na (~4562 kJ/mol) > IE_2 of Mg (~1451 kJ/mol)

After removing the first electron:

Na^+ has configuration [Ne] — the core noble-gas configuration (no valence electrons).

The second electron must be removed from the very stable, fully filled $2p^6$ shell (inner core).

This requires an enormous amount of energy.

Mg^+ has configuration [Ne] $3s^1$ — still has one 3s valence electron.

Removing this 3s valence electron is much easier than breaking into a noble-gas core.

Therefore, Na's second ionization enthalpy is much higher than Mg's second ionization enthalpy.

Answer: IE_1: Mg > Na because Mg has higher Z with both outer electrons in 3s. IE_2: Na >> Mg because Na^+ has a noble-gas core (removing from inner shell is extremely difficult), while Mg^+ still has a 3s valence electron that is relatively easy to remove.

Q18 What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?

SOLUTION

The ionization enthalpy decreases down a group because of the following factors:

1. Increase in atomic size (principal quantum number n increases):
   As we go down a group, electrons enter higher principal shells (larger n).
   These electrons are farther from the nucleus and are easier to remove.

2. Increase in shielding (screening) effect:
   More core electron shells intervene between the nucleus and the valence electron.
   These inner electrons shield the valence electron from the full nuclear charge.
   The effective nuclear charge experienced by the valence electron is reduced.

3. Increase in nuclear charge is offset by shielding:
   Although Z increases down the group, the increase in shielding by inner electrons is nearly as large or larger.
   The net effective nuclear charge on valence electrons does not increase significantly, or may even decrease.

Combined result of these factors:
Valence electrons are farther from the nucleus and more shielded => they experience a lower effective nuclear charge.
Therefore, less energy is required to remove the outermost electron => ionization enthalpy decreases down a group.

Answer: Ionization enthalpy decreases down a group because: (1) atomic radius increases (valence electrons farther from nucleus), (2) shielding by core electrons increases (reducing effective nuclear charge on valence electrons), so less energy is needed to remove the outermost electron.

Q19 The first ionization enthalpy values (in kJ mol^{-1}) of group 13 elements are: B (801), Al (577), Ga (579), In (558), Tl (589). How would you explain this deviation from the general trend?

SOLUTION

General expected trend: Ionization enthalpy should decrease from B to Tl as we go down the group.
Observed: B(801) > Al(577) < Ga(579) > In(558) < Tl(589)

Two anomalies: Ga > Al and Tl > In

Anomaly 1 — Ga has higher IE_1 than Al:

Ga (Z=31) comes after the 3d transition series (Sc to Zn, Z = 21-30).

The 3d electrons (10 electrons added before Ga) are poor shielders of nuclear charge.

d electrons are diffuse and do not effectively shield the outer 4s and 4p electrons from the nucleus.

As a result, the effective nuclear charge on Ga's valence electrons is higher than expected.

Ga is also slightly smaller than expected (the d-block contraction).

Therefore, Ga's IE_1 is slightly higher than Al's, breaking the expected downward trend.

Anomaly 2 — Tl has higher IE_1 than In:

Tl (Z=81) comes after both the 4f (lanthanide) series (14 elements) and the 5d transition series (10 elements).

The 4f electrons (lanthanide contraction) are very poor shielders — they cause the valence electrons to experience a much higher effective nuclear charge.

Additionally, relativistic effects become significant for heavy elements: the 6s electrons in Tl contract and are stabilized (relativistic contraction).

These combined effects make Tl's valence electrons harder to remove than In's.

Hence Tl > In in ionization enthalpy.

Answer: Ga > Al: Ga follows the 3d transition series; poor shielding by 3d electrons raises effective Z on Ga's valence electrons. Tl > In: Tl follows 4f (lanthanide contraction) and 5d series; poor shielding by f and d electrons plus relativistic stabilization of the 6s orbital raises effective Z on Tl.

Q20 Which of the following pairs of elements would have a more negative electron gain enthalpy? (i) O or F (ii) F or Cl

SOLUTION

(i) O or F:

F (fluorine) has a more negative electron gain enthalpy than O.

F: Z=9, configuration [He] $2s^2$ $2p^5$ — needs 1 more electron to complete the octet.

High nuclear charge, small size => strong attraction for an extra electron.

Delta_eg H(F) = -328 kJ/mol

O: Z=8, configuration [He] $2s^2$ $2p^4$ — gaining one electron gives O^- ($2p^5$).

Lower nuclear charge than F.

Delta_eg H(O) = -141 kJ/mol

F has a more negative electron gain enthalpy than O.

(ii) F or Cl:

Cl has a more negative electron gain enthalpy than F.

Although F has higher electronegativity, its electron gain enthalpy is less negative than Cl.

Reason: F is very small (atomic radius ~ 64 pm). Adding an electron to the already compact 2p orbital

causes significant electron–electron repulsion in the small n=2 shell.

Cl is larger (atomic radius ~ 99 pm), with electrons in n=3 shell.

Less inter-electron repulsion when an electron is added to the more spacious 3p orbital.

Delta_eg H: Cl (-348 kJ/mol) < F (-328 kJ/mol) — Cl is more negative.

Therefore, Cl has a more negative electron gain enthalpy than F.

Answer: (i) F has a more negative electron gain enthalpy than O (higher Z, more stable after gaining electron). (ii) Cl has a more negative electron gain enthalpy than F (despite lower electronegativity, Cl's larger 3p orbital accommodates the added electron with less repulsion).

Q21 Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.

SOLUTION

First electron gain enthalpy of O:

$O(g) + e^- \rightarrow O^-(g)$; $\Delta_{eg}H_1 = -141$ kJ/mol (negative — energy released)

O has $2p^4$; gaining one electron gives O^- with $2p^5$.

This is thermodynamically favourable (electron attracted by the nucleus).

Second electron gain enthalpy of O:

$O^-(g) + e^- \rightarrow O^{2-}(g)$; $\Delta_{eg}H_2 = +780$ kJ/mol (positive — energy absorbed)

The second electron gain enthalpy is positive (endothermic).

Justification:

O^- already carries a net negative charge.

Adding a second electron means forcing another negatively charged electron onto an already negatively charged ion.

The electron–electron repulsion between the incoming electron and the existing negative charge of O^- outweighs the nuclear attraction.

Energy must be supplied to add the second electron against this repulsion.

Therefore, the second electron gain enthalpy is positive (endothermic), unlike the first which was exothermic.

Answer: The second electron gain enthalpy of O is positive (energy must be absorbed). Adding an electron to the already negatively charged O^- ion requires overcoming the electrostatic repulsion between the incoming electron and the negative ion; this repulsion exceeds the nuclear attraction.

Q22 What is the basic difference between the terms electron gain enthalpy and electronegativity?

SOLUTION

Electron gain enthalpy:

Defined as the enthalpy change when a neutral gaseous atom gains an electron to form a gaseous anion:

$A(g) + e^- \rightarrow A^-(g);\ \Delta_{eg} H$

It is a precisely measurable thermodynamic quantity with units of kJ/mol.

It refers to an isolated, gaseous atom in the ground state.

It measures the energy change for a specific, defined process.

Electronegativity:

Defined as the ability of an atom in a covalent bond to attract the shared electron pair towards itself.

It is a relative, dimensionless concept — a tendency, not an absolute energy.

It has no SI units; it is expressed on arbitrary scales (Pauling, Mulliken, Allred-Rochow, etc.).

It describes an atom's behaviour in a MOLECULE (in the context of a chemical bond), not an isolated atom.

It varies slightly with the oxidation state and bonding environment of the atom.

Key difference:

Electron gain enthalpy is an experimentally measured energy change for an isolated gaseous atom.

Electronegativity is a bonding concept — a relative measure of electron-attracting power within a molecule.

Both increase across a period and decrease down a group, but they are not numerically related.

Answer: Electron gain enthalpy is the measured enthalpy change (kJ/mol) for a gaseous atom gaining one electron in isolation. Electronegativity is a dimensionless, relative measure of an atom's ability to attract bonding electrons towards itself within a covalent bond; it is a bonding property, not a thermodynamic quantity for an isolated atom.

Q23 How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds?

SOLUTION

The statement is not entirely accurate.

While 3.0 is the commonly cited Pauling electronegativity value for nitrogen, electronegativity is not a fixed, invariant property of an atom.

It is an average or representative value that can vary with:

1. Oxidation state: Nitrogen in N^{3-} (nitride, low oxidation state) has a different electron-attracting tendency than N in NO_2 (+4) or N in HNO_3 (+5).
2. Hybridisation: An sp hybridised N (as in HCN) is more electronegative than sp^3 hybridised N (as in NH_3), because s-character increases the effective nuclear charge felt by the bonding electrons.
3. Chemical environment: The nature and electronegativity of bonded atoms affect the electron density on N.

Therefore, one should treat 3.0 as an average or approximate value of nitrogen's electronegativity.

Pauling's scale assigns a single value per element as a useful reference, but in practice electronegativity is a context-dependent quantity.

The statement oversimplifies by claiming the value is exactly 3.0 in ALL nitrogen compounds.

Answer: The statement is an oversimplification. The value 3.0 is an average Pauling electronegativity for N. In reality, electronegativity varies with oxidation state, hybridisation (sp N is more electronegative than sp^3 N), and bonding environment. It is a useful comparative number, not a fixed property.

Q24 Describe the theory associated with the radius of an atom as it (a) gains an electron (b) loses an electron

SOLUTION

(a) When an atom gains an electron (forms an anion):

The nuclear charge (Z) remains unchanged.

The number of electrons increases by 1.

With the same nuclear charge attracting more electrons:

- Electron–electron repulsion among the increased number of electrons increases.

- The effective nuclear charge per electron decreases.

The electron cloud expands (electron–electron repulsion pushes electrons outward).

The atom is effectively shielded more per electron from the nucleus.

Result: the ionic radius of the anion is LARGER than the parent atom's radius.

Example: Cl (99 pm) -> Cl^- (181 pm)

(b) When an atom loses an electron (forms a cation):

The nuclear charge (Z) remains unchanged.

The number of electrons decreases by 1.

With the same nuclear charge now attracting fewer electrons:

- Electron–electron repulsion decreases.

- The effective nuclear charge per electron increases.

The remaining electrons are pulled closer to the nucleus.

If the outermost shell is completely emptied, a new, smaller shell becomes the outermost.

Result: the ionic radius of the cation is SMALLER than the parent atom's radius.

Example: Na (186 pm) -> Na^+ (102 pm)

Answer: (a) Anion (gains e^-): radius increases — same Z attracts more electrons, increasing e-e repulsion and reducing effective nuclear charge per electron, expanding the electron cloud. (b) Cation (loses e^-): radius decreases — same Z holds fewer electrons more tightly (higher effective nuclear charge per electron), contracting the electron cloud.

Q25 Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

SOLUTION

The first ionization enthalpies for two isotopes of the same element would be essentially the SAME.

Justification:

Isotopes of an element have the same atomic number (same number of protons) and the same number of electrons.

They differ only in the number of neutrons in the nucleus.

The ionization enthalpy depends on the electronic configuration of the atom — specifically on the effective nuclear charge experienced by the outermost electron, the principal quantum number n, and the shielding by inner electrons.

Since neutrons are neutral particles, they do not directly affect the electrostatic interaction between the nucleus and the electrons.

The electronic configuration, effective nuclear charge, and energy levels are the same for all isotopes of an element.

Therefore, the energy required to remove the outermost electron is virtually identical for isotopes.

Any extremely tiny difference (due to slightly different nuclear mass affecting the Bohr energy levels — the isotope shift) is negligibly small in practice.

Answer: The first ionization enthalpies of isotopes of the same element are essentially the same. Isotopes have the same Z (same number of electrons and same electronic configuration); neutrons are neutral and do not affect the electron–nucleus electrostatic interaction that determines ionization enthalpy.

Q26 What are the major differences between metals and non-metals?

SOLUTION

Major differences between metals and non-metals:

Physical properties:

Metals: lustrous (shiny), good conductors of heat and electricity, malleable, ductile, high melting/boiling points (generally), solid at room temperature (except Hg).

Non-metals: generally dull, poor conductors (insulators or semiconductors),

brittle (if solid), low melting/boiling points, exist as solids, liquids, or gases at room temperature.

Chemical properties:

Metals: lose electrons (electropositivity), form cations, have low ionization enthalpies, low electronegativity.

Non-metals: gain electrons (electronegativity), form anions or share electrons, have high ionization enthalpies, high electronegativity.

Oxides:

Metals: form basic oxides (e.g., Na_2O, CaO) that react with water to form bases.

Non-metals: form acidic oxides (e.g., SO_2, CO_2, P_4O_{10}) that react with water to form acids.

Position in periodic table:

Metals: generally on the left and centre of the periodic table (s-block, d-block, some p-block).

Non-metals: generally on the upper right of the periodic table (p-block).

Metalloids (semi-metals) lie at the dividing line.

Answer: Metals: lustrous, malleable, ductile, good conductors, low ionization enthalpy, low electronegativity; lose electrons to form cations; form basic oxides. Non-metals: dull (if solid), poor conductors, brittle, high ionization enthalpy, high electronegativity; gain/share electrons; form acidic oxides.

Q27 Use the periodic table to answer the following questions. (a) Identify an element with five electrons in the outer subshell. (b) Identify an element that would tend to lose two electrons. (c) Identify an element that would tend to gain two electrons. (d) Identify the group having metal, non-metal, liquid as well as gas at the room temperature.

SOLUTION

(a) Element with 5 electrons in the outer subshell:

Outer subshell with 5 electrons: could be p^5 or d^5.

p^5: Group 17 (halogens) — e.g., Fluorine (F) or Chlorine (Cl).

d^5: Group 7 (Mn) or mid-d-block elements (e.g., Mn: [Ar] $3d^5\ 4s^2$).

Most commonly expected answer: any halogen (Group 17, $ns^2\ np^5$).

Examples: F (Z=9), Cl (Z=17)

(b) Element tending to lose two electrons:

Elements with 2 valence electrons that they easily lose => Group 2 (alkaline earth metals).

Examples: Mg (Z=12, [Ne] $3s^2$ -> Mg^{2+}), Ca (Z=20, [Ar] $4s^2$ -> Ca^{2+})

(c) Element tending to gain two electrons:

Elements needing 2 electrons to complete their octet => Group 16 (chalcogens, $ns^2\ np^4$).

Examples: O (Z=8, gains $2e^-$ -> O^{2-}), S (Z=16, gains $2e^-$ -> S^{2-})

(d) Group having metal, non-metal, liquid, and gas at room temperature:

Group 17 (halogens) satisfies all four conditions:

- F_2: gas at room temperature (non-metal)
- Cl_2: gas at room temperature (non-metal)
- Br_2: liquid at room temperature (non-metal)
- I_2: solid at room temperature (non-metal)
- At (astatine): has metallic character

Thus Group 17 contains gases (F_2, Cl_2), a liquid (Br_2), a solid non-metal (I_2), and At with metallic character — satisfying all four states/types.

Answer: (a) Group 17 halogens (e.g., F or Cl, $ns^2\ np^5$ = 5 electrons in outer p subshell). (b) Group 2 elements (e.g., Mg, Ca). (c) Group 16 elements (e.g., O, S). (d) Group 17 (halogens): F_2 and Cl_2 are gases, Br_2 is a liquid, I_2 is a solid non-metal, and At has metallic character.

Q28 The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that among group 17 elements is F > Cl > Br > I. Explain.

SOLUTION

Group 1 (alkali metals) — reactivity increases down the group (Li < Na < K < Rb < Cs):

Alkali metals react by losing their single valence electron.

Down the group:

- Atomic radius increases (electrons in higher n shells, farther from nucleus).
- More shielding by core electrons reduces effective nuclear charge on valence electron.
- Ionization enthalpy decreases progressively.
- The valence electron is increasingly easier to remove.

Therefore, the tendency to lose an electron (and hence reactivity) increases: Li <

Na < K < Rb < Cs.

Group 17 (halogens) — reactivity decreases down the group (F > Cl > Br > I):

Halogens react by gaining an electron (oxidising agents).

Down the group:

- Atomic radius increases.
- Electron gain enthalpy becomes less negative (atom less eager to accept an electron).
- Electronegativity decreases.
- The tendency to gain an electron decreases.

Additionally, F is unique: it has the highest electronegativity, it cannot expand its octet (no d-orbitals in n=2), and it forms only F^- — it is always the oxidising agent.

Therefore, oxidising power and reactivity decrease: F > Cl > Br > I.

Answer: Group 1: reactivity increases down the group because ionization enthalpy decreases (larger atoms, more shielding), making it progressively easier to lose one electron. Group 17: reactivity decreases down the group because electron gain enthalpy becomes less negative (larger atoms accommodate added electron less strongly), reducing the tendency to gain an electron.

Q29 Write the general outer electronic configuration of s-, p-, d- and f-block elements.

SOLUTION

s-block elements (Groups 1 and 2):

General outer electronic configuration: ns^{1-2}

(Group 1: ns^1; Group 2: ns^2; n = 1 to 7)

p-block elements (Groups 13 to 18):

General outer electronic configuration: $ns^2\ np^{1-6}$

(The number of p electrons ranges from 1 in Group 13 to 6 in Group 18; n = 2 to 6)

d-block elements (Groups 3 to 12, transition elements):

General outer electronic configuration: $(n-1)d^{1-10}\ ns^{1-2}$

(The d subshell being filled is one principal quantum number below the outermost s subshell; n = 4 to 7)

f-block elements (lanthanides and actinides):

General outer electronic configuration: $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$

(The f subshell being filled is two principal quantum numbers below the outermost s subshell; n = 6 for lanthanides, n = 7 for actinides)

Answer: s-block: ns^{1-2}; p-block: $ns^2\ np^{1-6}$; d-block: $(n-1)d^{1-10}\ ns^{1-2}$; f-block: $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$.

Q30 Assign the position of the element having outer electronic configuration (i) $ns^2\ np^4$ for n=3, (ii) $(n-1)d^2\ ns^2$ for n=4, and (iii) $(n-2)f^7\ (n-1)d^1\ ns^2$ for n=6, in the periodic table.

SOLUTION

(i) $ns^2\ np^4$ for n = 3:
Configuration: $3s^2\ 3p^4$
Period: n = 3 => Period 3
Block: p-block (last subshell = 3p)
Group: Total valence electrons = 2 + 4 = 6 => Group 16
Element: Sulfur (S), Z = 16

(ii) $(n-1)d^2\ ns^2$ for n = 4:
Configuration: $3d^2\ 4s^2$
Period: n = 4 => Period 4
Block: d-block (last subshell = 3d)
Group: For d-block, group number = number of d electrons + number of s electrons = 2 + 2 = 4 => Group 4
Element: Titanium (Ti), Z = 22

(iii) $(n-2)f^7\ (n-1)d^1\ ns^2$ for n = 6:
Configuration: $4f^7\ 5d^1\ 6s^2$
Period: n = 6 => Period 6
Block: f-block (lanthanides; 4f is being filled)
Group: f-block elements are placed in Group 3 (IUPAC places lanthanides at Group 3 position in Period 6)
Element: Gadolinium (Gd), Z = 64

Answer: (i) Period 3, Group 16, p-block — Sulfur (S, Z=16). (ii) Period 4, Group 4, d-block — Titanium (Ti, Z=22). (iii) Period 6, Group 3 (lanthanide series), f-block — Gadolinium (Gd, Z=64).

Q31 The first (ΔH_1) and the second (ΔH_2) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg}H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:

Element I: ΔH_1=520, ΔH_2=7300, $\Delta_{eg}H$=-60

Element II: ΔH_1=419, ΔH_2=3051, $\Delta_{eg}H$=-48

Element III: ΔH_1=1681, ΔH_2=3374, $\Delta_{eg}H$=-328

Element IV: ΔH_1=1008, ΔH_2=1846, $\Delta_{eg}H$=-295

Element V: ΔH_1=2372, ΔH_2=5251, $\Delta_{eg}H$=+48

Element VI: ΔH_1=738, ΔH_2=1451, $\Delta_{eg}H$=-40

Which of the above elements is likely to be: (a) the least reactive element, (b) the most reactive metal, (c) the most reactive non-metal, (d) the least reactive non-metal, (e) the metal which can form a stable binary halide of the formula MX_2 (X=halogen), (f) the metal which can form a predominantly stable covalent halide of the formula MX (X=halogen)?

SOLUTION

Analysis of each element:

Element I (IE_1=520, IE_2=7300, EGE=-60):

Huge jump from IE_1 to IE_2 (factor ~14) => loses 1 electron easily, hard to lose 2nd.

=> Group 1 alkali metal (similar to Na: IE_1~496, IE_2~4562). Moderately reactive metal.

Element II (IE_1=419, IE_2=3051, EGE=-48):

Huge jump from IE_1 to IE_2 (factor ~7.3) => loses 1 electron very easily.

Lowest IE_1 of all => most electropositive metal. (Similar to K: IE_1=419 kJ/mol).

=> Most reactive metal (Group 1).

Element III (IE_1=1681, IE_2=3374, EGE=-328):

Very high IE_1, very negative EGE => highly electronegative non-metal.

IE_1=1681 matches F; EGE=-328 matches F.

=> Most reactive non-metal (Fluorine, Group 17).

Element IV (IE_1=1008, IE_2=1846, EGE=-295):

High IE_1 and 2 (both moderately high), negative EGE => non-metal.

Less reactive than element III; negative EGE shows tendency to gain electrons.

=> Least reactive non-metal in this list (less reactive than F, similar to Cl or S).

Element V (IE_1=2372, IE_2=5251, EGE=+48):

Very high IE_1, positive EGE (does not gain electrons) => does not easily gain or

lose electrons.

Positive EGE uniquely identifies a noble gas.

=> Least reactive element (noble gas).

Element VI (IE_1=738, IE_2=1451, EGE=-40):

Moderate jump from IE_1 to IE_2 (factor ~2) => loses 2 electrons readily.

IE_1~738 matches Mg (737.7 kJ/mol); loses 2 electrons to form M^{2+}.

=> Group 2 metal; forms stable binary halide MX_2 (e.g., $MgCl_2$).

Answers:

(a) Least reactive element: V (noble gas, positive EGE, very high IE)

(b) Most reactive metal: II (lowest IE_1, loses 1 electron most easily, Group 1 metal like K)

(c) Most reactive non-metal: III (highest negative EGE=-328, high IE_1; F-like, Group 17)

(d) Least reactive non-metal: IV (less electronegative than III; gains electrons but less readily)

(e) Metal forming stable MX_2: VI (Group 2, Mg-like, loses 2 electrons => MgX_2-type halide)

(f) Metal forming predominantly covalent MX: I (Group 1 metal like Na; forms monohalide NaX; high IE_2 confirms +1 only; small size increases polarizing power toward covalent character)

Answer: (a) V — noble gas (highest IE, positive EGE). (b) II — most reactive metal (lowest IE_1, Group 1 like K). (c) III — most reactive non-metal (EGE = -328, F-like). (d) IV — least reactive non-metal. (e) VI — forms MX_2 (Group 2, Mg-like). (f) I — forms predominantly covalent MX (Group 1 metal with high polarizing power).

Q32 Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements. (a) Lithium and oxygen (b) Magnesium and nitrogen (c) Aluminium and iodine (d) Silicon and oxygen (e) Phosphorus and fluorine (f) Element 71 and fluorine

SOLUTION

Determine oxidation states from group positions, then balance charges:

(a) Lithium (Li, Group 1, +1) and Oxygen (O, Group 16, -2):

Cross-multiply charges: 2 Li^+ and 1 O^{2-} => Li_2O

Formula: Li_2O (lithium oxide)

Charge check: 2(+1) + (-2) = 0 ✓

(b) Magnesium (Mg, Group 2, +2) and Nitrogen (N, Group 15, -3):
Cross-multiply charges: 3 Mg^{2+} and 2 N^{3-} => Mg_3N_2
Formula: Mg_3N_2 (magnesium nitride)
Charge check: 3(+2) + 2(-3) = 0 ✓

(c) Aluminium (Al, Group 13, +3) and Iodine (I, Group 17, -1):
Cross-multiply charges: 1 Al^{3+} and 3 I^- => AlI_3
Formula: AlI_3 (aluminium iodide)
Charge check: +3 + 3(-1) = 0 ✓

(d) Silicon (Si, Group 14, +4) and Oxygen (O, Group 16, -2):
Cross-multiply charges: 1 Si^{4+} and 2 O^{2-} => SiO_2
Formula: SiO_2 (silicon dioxide)
Charge check: +4 + 2(-2) = 0 ✓

(e) Phosphorus (P, Group 15) and Fluorine (F, Group 17, -1):
P can exhibit +3 or +5 oxidation state (can use $3p^3$ or promote electrons to 3d, using all 5 valence electrons).
Both PF_3 (P is +3) and PF_5 (P is +5) are stable compounds.
Formulas: PF_3 (phosphorus trifluoride) and PF_5 (phosphorus pentafluoride)
Charge check PF_3: +3 + 3(-1) = 0 ✓; PF_5: +5 + 5(-1) = 0 ✓

(f) Element 71 (Lutetium, Lu, lanthanide, +3) and Fluorine (F, -1):
Cross-multiply charges: 1 Lu^{3+} and 3 F^- => LuF_3
Formula: LuF_3 (lutetium fluoride)
Charge check: +3 + 3(-1) = 0 ✓

Answer: (a) Li_2O, (b) Mg_3N_2, (c) AlI_3, (d) SiO_2, (e) PF_3 and PF_5, (f) LuF_3.

Q33 In the modern periodic table, the period indicates the value of: (a) atomic number (b) atomic mass (c) principal quantum number (d) azimuthal quantum number.

SOLUTION

The period number in the modern periodic table corresponds to the principal quantum number (n) of the outermost shell (highest energy shell) of the elements in that period.

For example:

Period 1: outermost shell n = 1 (H, He)
Period 2: outermost shell n = 2 (Li to Ne)
Period 3: outermost shell n = 3 (Na to Ar)
Period 4: outermost shell n = 4 (K to Kr)

Atomic number and mass are not indicated by the period directly.
The azimuthal quantum number (l) indicates the block (s, p, d, f), not the period.

Answer: (c) principal quantum number

Answer: (c) principal quantum number. The period number equals the principal quantum number n of the outermost electron shell.

Q34 Which of the following statements related to the modern periodic table is incorrect? (a) The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-shell. (b) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell. (c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell. (d) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

SOLUTION

Evaluate each statement:

(a) p-block has 6 columns: TRUE.
A p subshell has 3 orbitals (m_l = -1, 0, +1) x 2 spins = 6 electrons maximum.
Therefore p-block spans Groups 13 to 18 = 6 columns.

(b) d-block has 8 columns: INCORRECT (FALSE).
A d subshell has 5 orbitals (m_l = -2, -1, 0, +1, +2) x 2 spins = 10 electrons maximum.
Therefore d-block should have 10 columns (Groups 3 to 12), NOT 8.
This statement is the incorrect one.

(c) Each block has columns equal to electrons that can fill its subshell: TRUE.
s-block: 2 electrons, 2 columns. p-block: 6 electrons, 6 columns.
d-block: 10 electrons, 10 columns. f-block: 14 electrons, 14 columns.

(d) Block indicates l value for last subshell receiving electrons: TRUE.
s-block: l = 0; p-block: l = 1; d-block: l = 2; f-block: l = 3.

The INCORRECT statement is (b).

Answer: (b) is incorrect. The d-block has 10 columns (not 8), because a d subshell can hold a maximum of 10 electrons (5 orbitals x 2 electrons each).

Q35 Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell? (a) Valence principal quantum number (n) (b) Nuclear charge (Z) (c) Nuclear mass (d) Number of core electrons.

SOLUTION

Evaluate how each factor affects valence electrons:

(a) Valence principal quantum number (n): AFFECTS the valence shell.
Determines the energy level of the valence electrons and the size of the valence orbital.

(b) Nuclear charge (Z): AFFECTS the valence shell.
Higher Z increases the attractive force on all electrons including valence electrons.
It directly governs ionization enthalpy and electronegativity.

(c) Nuclear mass: does NOT significantly affect the valence shell.
The nuclear mass (number of neutrons) does not change the nuclear charge.
Neutrons are electrically neutral and do not exert electromagnetic force on electrons.
The energy levels and orbitals of electrons depend on Z (charge), not the mass of the nucleus.
(Isotope effects on energy levels are extremely small and negligible for chemistry.)

(d) Number of core electrons: AFFECTS the valence shell.
Core electrons shield the valence electrons from the full nuclear charge.
The shielding (screening) effect determines the effective nuclear charge Z_eff felt by valence electrons.

The factor that does NOT affect the valence shell is (c) Nuclear mass.

Answer: (c) Nuclear mass does not affect the valence shell. Neutrons are electrically neutral and do not exert electromagnetic forces on electrons; only nuclear charge (Z), n, and core electron shielding affect the valence shell.

Q36 The size of isoelectronic species — F^-, Ne and Na^+ is affected by (a) nuclear charge (Z) (b) valence principal quantum number (n) (c) electron-electron interaction in the outer orbitals (d) none of the factors because their size is the same.

SOLUTION

F^-, Ne and Na^+ are isoelectronic: all have 10 electrons with configuration $1s^2\ 2s^2\ 2p^6$.

Their nuclear charges are different:

F^-: Z = 9

Ne: Z = 10

Na^+: Z = 11

For isoelectronic species (same number of electrons, same principal quantum number n = 2 for valence electrons):

- The value of n is the same for all three => (b) n alone cannot explain the size difference.

- The electron–electron repulsion is similar for all three (same configuration, same number of e^-) => (c) is not the differentiating factor.

- The nuclear charge Z is different for each:

Higher Z = more protons pulling the same 10 electrons closer = smaller radius.

$Na^+(Z=11) < Ne(Z=10) < F^-(Z=9)$ in order of increasing size.

Therefore, the size of isoelectronic species is primarily determined by the nuclear charge Z.

Answer: (a) nuclear charge (Z)

Answer: (a) Nuclear charge (Z). For isoelectronic species with the same number of electrons and same n, the higher the nuclear charge, the more strongly the electrons are pulled inward: Na^+ (Z=11) < Ne (Z=10) < F^- (Z=9) in size.

Q37 Which one of the following statements is incorrect in relation to ionization enthalpy? (a) Ionization enthalpy increases for each successive electron. (b) The greatest increase in ionization enthalpy is experienced on removal of electron from core noble gas configuration. (c) End of valence electrons is marked by a big jump in ionization enthalpy. (d) Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.

SOLUTION

Evaluate each statement:

(a) Ionization enthalpy increases for each successive electron: TRUE.

As each electron is removed, the cation has increasing positive charge, which attracts the remaining electrons more strongly.

Therefore, successive ionization enthalpies always increase.

(b) Greatest increase in ionization enthalpy is on removal from core noble-gas configuration: TRUE.

After all valence electrons are removed, the next electron must come from the noble-gas core.

Noble-gas electron configurations are extremely stable — a huge amount of energy is needed.

This is the 'big jump' used to identify group membership.

(c) End of valence electrons is marked by a big jump in ionization enthalpy: TRUE.

This is consistent with statement (b) and is the basis for determining valence electrons.

(d) Removal of electron from orbitals bearing lower n value is easier than from higher n value: INCORRECT (FALSE).

Electrons in orbitals with LOWER n are closer to the nucleus and more strongly bound.

Lower n => higher binding energy => harder to remove.

Electrons in orbitals with HIGHER n (outer shells) are farther from the nucleus => easier to remove.

Therefore, the statement has it backwards: removal from HIGHER n is easier.

The INCORRECT statement is (d).

Answer: (d) is incorrect. Electrons in orbitals with lower n (inner shells) are MORE tightly bound (harder to remove), not easier. It is the electrons in higher n orbitals (outer shells) that are removed more easily.

Q38 Considering the elements B, Al, Mg, and K, the correct order of their metallic character is: (a) B > Al > Mg > K (b) Al > Mg > B > K (c) Mg > Al > K > B (d) K > Mg > Al > B

SOLUTION

Metallic character = tendency to lose electrons (electropositivity).

Metallic character increases down a group and decreases across a period (left to right).

Positions in the periodic table:

B (Boron): Period 2, Group 13 — metalloid, very low metallic character

Al (Aluminium): Period 3, Group 13 — metal, but less metallic than Mg

Mg (Magnesium): Period 3, Group 2 — metal, more metallic than Al (same period, Al is to the right)

K (Potassium): Period 4, Group 1 — alkali metal, highest metallic character

Decreasing metallic character: K > Mg > Al > B

Reasoning:

K (Period 4, Group 1): very low IE_1, loses 1 electron very easily — most metallic.

Mg (Period 3, Group 2): loses 2 electrons; more metallic than Al (lower IE, left of Al in same period).

Al (Period 3, Group 13): loses 3 electrons; less metallic than Mg, more than B.

B (Period 2, Group 13): high IE for its group position; metalloid — least metallic.

Answer: (d) K > Mg > Al > B. Metallic character increases down a group and decreases across a period; K (Group 1, Period 4) is most metallic and B (Period 2 metalloid) is least.

Q39 Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is: (a) B > C > Si > N > F (b) Si > C > B > N > F (c) F > N > C > B > Si (d) F > N > C > Si > B

SOLUTION

Non-metallic character = tendency to gain electrons = increases with increasing electronegativity.

Non-metallic character increases across a period (left to right) and decreases down a group.

Positions:

Si: Period 3, Group 14 — metalloid, least non-metallic in this set

B: Period 2, Group 13 — metalloid, slightly more non-metallic than Si

C: Period 2, Group 14 — non-metal

N: Period 2, Group 15 — non-metal, more electronegative than C

F: Period 2, Group 17 — most electronegative element, highest non-metallic character

Order (increasing non-metallic character): Si < B < C < N < F

Decreasing order: F > N > C > B > Si

This matches option (c): F > N > C > B > Si

Answer: (c) F > N > C > B > Si. Non-metallic character increases across Period 2 (B < C < N < F) and is lower for Period 3 Si (metalloid) than Period 2 B.

Q40 Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidizing property is: (a) F > Cl > O > N (b) F > O > Cl > N (c) Cl > F > O > N (d) O > F > N > Cl

SOLUTION

Oxidizing power = tendency to gain electrons = increases with electronegativity and electron gain enthalpy.

Electronegativity (Pauling scale):

F = 4.0 (highest of all elements)

O = 3.5

Cl = 3.0

N = 3.0

F is the strongest oxidising agent of all elements — it always acts as the oxidiser.

O is the second most electronegative element and a powerful oxidiser.
Cl: although Cl is a strong oxidiser, its electronegativity (3.0) is less than O (3.5).
N: electronegativity 3.0 (same as Cl), but N^2 is very stable (triple bond with high bond energy = 945 kJ/mol), making it a poor oxidiser under normal conditions.

Oxidising power order: F > O > Cl > N

Note: Although Cl has a more negative electron gain enthalpy than O (Cl: -348 vs O: -141 kJ/mol), O is a stronger oxidising agent in most reactions because of its higher electronegativity (3.5 vs 3.0) and the very high energy of its element form (which drives oxide formation).
F > O > Cl > N corresponds to option (b).

Answer: (b) F > O > Cl > N. F is the strongest oxidising agent (highest electronegativity, 4.0); O (3.5) is a stronger oxidiser than Cl (3.0) despite Cl's more negative electron gain enthalpy; N_2 is very stable and the weakest oxidiser here.