تغطية اختبار نهاية الفصل





تم تحميل هذا الملف من موقع المناهج الإماراتية	
موقع المناهج ⇦ المناهج الإماراتية ⇦ الصف العاشر المتقدم ⇦ كيمياء ⇦ الفصل الأول ⇦ اختبارات ⇦ الملف	
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التواصل الاجتماعي بحسب الصف العاشر المتقدم					
صفحة المناهج الإماراتية على فيسبوك					
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المزيد من الملفات بحسب الصف العاشر المتقدم والمادة كيمياء في الفصل الأول	
تغطية اختبار نهاية الفصل	1
مراجعة هيكل اختبار نهاية الفصل	2
اختبار نهاية الفصل (محاكاة إلكترونية - بريدج)	3
شرح نموذج بور والمدارات الكمية وسلاسل الطيف الكيميائي – انسباير	4
حل مراجعة هيكل : البناء الذري ومستويات الطاقة	5

تحميل تطبيق المناهج للهواتف الذكية	
تطبيق المناهج على Store App (آبل)	تطبيق المناهج على Play Google (أندرويد)

Grade 10 Advanced

Chemistry

End-of-Term 1 Exam – Academic Year 2026 – 2027

Item	Learning outcome	Reference
Q1	Differentiate between the ground and excited states of an atom	p. 13 – Figure 10
Q2	Relate the principal quantum number to the total number of orbitals of the principal energy levels and the orbitals in sublevels	p. 22 – Table 2
Q3	Explain the Aufbau principle, the Pauli exclusion principle, and Hund's rule	pp. 23–24 – Figure 18 + Table 3
Q4	Write the electron configuration, orbital diagram, and noble gas notation of different elements	pp. 25–26 – Table 5
Q5	Represent the electron dot structure of atoms	pp. 28–29 – Table 6 + Example 3
Q6	List the families of the periodic table with the group number and the elements	pp. 40–41
Q7	Differentiate between representative elements and transition elements	p. 40
Q8	Use the electron configuration, orbital and noble gas notation of an element (Z 1–36) to locate it (period, group, block)	pp. 44–48 – Example 1 + Figure 9
Q9	The electron configuration: filling of sublevels	p. 46
Q10	Explain the trend of first ionization energy across a period (exceptions between groups 2 & 3 and 5 & 6) and down a group	p. 54 – Figure 17
Q11	Predict the periodic trend of electronegativity across a period and down a group	pp. 56–57 – Figure 19
Q12	Describe how ions (cations and anions) form to fulfill the octet rule	p. 65
Q13	Describe how ions form to fulfill the octet rule – pseudo noble gas configuration	p. 66 – Figure 3
Q14	Ion formation	p. 66 – Figure 4
Q15	Describe how ions (cations and anions) form to fulfill the octet rule	p. 69 – Table 4
Q16	Explain the physical properties of ionic compounds: melting and boiling point, conductivity (solid, molten, aqueous) and solubility	pp. 72–73 – Table 5 + Figure 8
Q17	Explain the physical properties of ionic compounds: melting and boiling point, conductivity and solubility	p. 73
Q18	Write the chemical formula (stock method) of an ionic compound containing monoatomic and polyatomic ions	pp. 75–78 – Tables 7, 8 + Examples 1, 2
Q19	Write the chemical formula (stock method) of an ionic compound containing monoatomic and polyatomic ions	pp. 79–80 – Tables 10, 11
Q20	Describe the formation of metallic bonds	p. 82 – Figure 10
Q21	Explain some physical properties of metals	pp. 83–84 – Figure 11 + Table 12
Q22	Define a covalent bond using an energy diagram	p. 92 – Figure 2
Q23	Differentiate between sigma and pi bonds	pp. 96–97 – Figures 7, 8, 9
Q24	Identify the relationship between the type of covalent bond (single, double, triple) and bond length, bond strength and dissociation energy	p. 98 – Table 1
Q25	Name a binary molecular compound from its molecular formula	pp. 100–101 – Table 2 + Example 2
Q26	Determine the chemical formula of a compound from its name	pp. 102–104 – Table 3 + Figure 12
Q27	Determine the chemical formula of a compound from its name	p. 103 – Table 4
Q28	Draw Lewis structures for covalent compounds with single and multiple bonds	pp. 107–110 – Examples 3, 4, 5
Q29	Identify resonance structures (O3, NO3−, NO2−, SO3 2−, CO3 2−)	pp. 111–112 – Figure 14
Q30	Determine exceptions to the octet rule (odd number of valence electrons, sub-octets, coordinate covalent bonds, expanded octets)	pp. 112–114 – Figures 16, 17 + Example 6

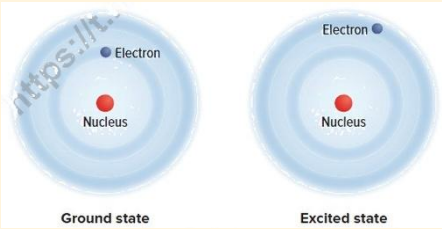
Note: questions may appear in a different order in the actual exam. Questions are arranged by the learning outcomes; under each outcome you will find the textbook reference figure, followed by practice questions.

Q1 — Learning outcome: Differentiate between the ground and excited states of an atom

Textbook: p. 13 – Figure 10

Key points from the lesson

- The ground state is the lowest allowable energy state of an atom.
- When an atom gains energy, its electron moves to a higher orbit (energy level): the atom is in an excited state.
- Bohr: the smaller the electron's orbit, the lower the atom's energy. Energy levels are quantized: n = 1, 2, 3 …
- When the electron drops back to a lower level, the atom emits a photon whose energy equals the difference between the two levels.



- Drops to n = 1 → Lyman (UV); to n = 2 → Balmer (visible); to n = 3 → Paschen (infrared).

Question 1 [Past exam 2025–26]

Which correctly describes an electron in a hydrogen atom according to Bohr's model?

A. It moves around the nucleus in only certain allowed circular orbits
B. It is stationary around the nucleus in only certain allowed circular orbits
C. It moves around the nucleus in unspecified circular orbits
D. It is stationary around the nucleus in unspecified circular orbits

Question 2 [Past exam 2025–26]

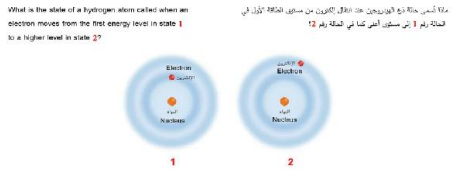
What explains the emission of the radiation series shown in the figure (hydrogen atom, levels n = 1 to 7)?



A. The electron drops from a higher energy level to a lower energy level
B. The electron moves from a lower energy level to a higher energy level
C. The electron moves around the nucleus at the same energy level
D. The hydrogen atom is in its ground state

Question 3 [Past exam 2025–26]

What is the state of a hydrogen atom called when its electron moves from the first energy level (state 1) to a higher level (state 2)?



A. Excited state
B. Ground state
C. Atomic state
D. Wave state

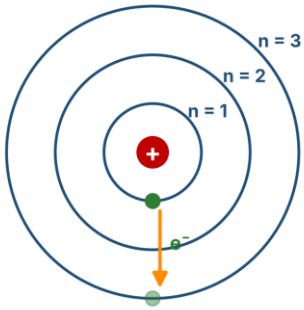
Question 4 [Mock exam]

In Bohr's model of the hydrogen atom, the electron moves around the nucleus in:

A. Circular orbits with fixed energy
B. Random paths with any energy
C. Inside the nucleus
D. A straight line away from the nucleus

Question 5 [Mock exam]

In the figure, the hydrogen electron moved from orbit n = 1 to orbit n = 3. The atom is now in the:



A. Ground state
B. Excited state
C. Stable state
D. Ionic state

Question 6 [Practice]

In Bohr's model, electrons are not allowed to occupy…

A. spaces between the orbits
B. any orbit
C. the nucleus
D. the ground state

Question 7 [Practice]

What is the ground state of the hydrogen atom in Bohr's model?

A. Electron in the orbit farthest from the nucleus
B. Electron in the orbit closest to the nucleus
C. Electron in any orbit
D. Electron outside the nucleus

Question 8 [Practice]

What did Bohr assign to each orbit in his model?

A. A color
B. A quantum number (n)
C. A mass
D. A charge

Question 9 [Practice]

What happens to electrons in Bohr's model when energy is added?

A. They move to different (higher) orbits
B. They stay in the same orbit
C. They disappear
D. They merge with the nucleus

Question 10 [Practice]

What is the ground state of an atom?

A. The lowest allowable energy state
B. The highest allowable energy state
C. An intermediate energy state
D. A state with no energy

Question 11 [Practice]

According to newer evidence, electrons around the nucleus…

A. move in circular orbits
B. move in elliptical orbits
C. move in fixed paths
D. do not move in circular orbits — only their probable location (cloud) can be known

Question 12 [Practice]

What happens when an atom gains energy?

A. It moves to a lower energy state
B. It remains in the ground state
C. It moves to an excited state
D. It loses electrons

Question 13 [Practice]

The lowest allowable energy state of an atom is called the:

A. excited state
B. ionized state
C. transition state
D. ground state

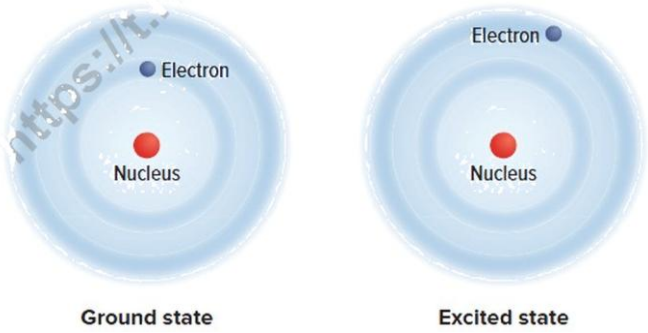
Question 14 [Practice]

An atom moves to an excited state when it:

A. absorbs energy
B. loses a proton
C. emits a photon
D. gains a neutron

Question 15 [Practice]

In Figure 10, compared with the ground state, the electron in the excited state is:



A. inside the nucleus
B. farther from the nucleus
C. removed from the atom
D. closer to the nucleus

Question 16 [Practice]

When an electron drops from n = 4 to n = 2 in hydrogen, the atom:

A. loses an electron
B. absorbs an infrared photon
C. emits an UV photon (Lyman)
D. emits a visible photon (Balmer)

Question 17 [Practice]

Which statement is correct?

A. Electrons can have any amount of energy in Bohr's model.
B. The ground state of hydrogen corresponds to n = 0.
C. The emitted photon energy equals the energy difference between the two levels.
D. An excited atom is more stable than a ground-state atom.

Q2

Learning outcome: Relate the principal quantum number to the total number of orbitals of the principal energy levels and the orbitals in sublevels

Textbook: p. 22 – Table 2

Table 2 Hydrogen's First Four Principal Energy Levels

Principal Quantum Number (n)	Sublevels (Types of Orbitals) Present	Number of Orbitals Related to Sublevel	Total Number of Orbitals Related to Principal Energy Level (n^2)
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

Key points from the lesson

- The principal quantum number (n) indicates the relative size and energy of atomic orbitals; energy increases as n increases.
- Principal energy level n contains n sublevels: n = 1 → s; n = 2 → s, p; n = 3 → s, p, d; n = 4 → s, p, d, f.
- Orbitals in each sublevel: s = 1, p = 3, d = 5, f = 7.
- Total orbitals in level n = n^2 (1, 4, 9, 16). Each orbital holds 2 electrons, so level n holds $2n^2$ electrons.
- 1s and 2s are both spherical; 2s is larger and higher in energy.

Question 1 [Past exam 2025–26]

How many energy sublevels does the n = 4 principal energy level have?

A. 4
B. 2
C. 3
D. 1

Question 2 [Mock exam]

The quantum number that describes the main energy level of an electron is:

A. Angular momentum quantum number (l)
B. Magnetic quantum number (ml)
C. Principal quantum number (n)
D. Spin quantum number (ms)

Question 3 [Mock exam]

How many orbitals are in the principal energy level n = 3?

A. 3
B. 6
C. 18
D. 9

Question 4 [Practice]

As n increases, the orbital…

A. becomes smaller
B. disappears
C. stays the same size
D. becomes larger

Question 5 [Practice]

The ‘math rule’ for the total number of orbitals in a main energy level is…

A. n
B. 2n
C. n^2
D. $2n^2$

Question 6 [Practice]

Using $2n^2$ for each level, what is the total maximum number of electrons that the first three principal energy levels can hold together?

A. 18
B. 28
C. 14
D. 36

Question 7 [Practice]

As n increases, the atom’s energy…

A. decreases
B. stays the same
C. increases
D. fluctuates

Question 8 [Practice]

The ‘math rule’ for the number of sublevels in a main energy level n is…

A. n
B. 2n
C. n^2
D. $2n^2$

Question 9 [Practice]

The lowest principal energy level is…

A. n = 0
B. n = 1
C. n = 2
D. n = 3

Question 10 [Practice]

The ‘math rule’ for the maximum number of electrons in a main energy level is…

A. n
B. 2n
C. n^2
D. $2n^2$

Question 11 [Practice]

How many orbitals are in an f sublevel?

A. 1
B. 3
C. 5
D. 7

Question 12 [Practice]

How many electrons are needed to completely fill the third energy level?

A. 3
B. 9
C. 18
D. 27

Question 13 [Practice]

Using Table 2, how many orbitals are in principal energy level n = 3?

Table 2 Hydrogen's First Four Principal Energy Levels

Principal Quantum Number (n)	Sublevels (Types of Orbitals) Present	Number of Orbitals Related to Sublevel	Total Number of Orbitals Related to Principal Energy Level (n^2)
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

A. 3
B. 6
C. 9
D. 18

Question 14 [Practice]

How many orbitals are in a d sublevel?

A. 1
B. 3
C. 5
D. 7

Question 15 [Practice]

Which type of sublevel appears for the first time in n = 3?

A. p
B. d
C. s
D. f

Question 16 [Practice]

What is the maximum number of electrons in energy level n = 4?

A. 8
B. 16
C. 18
D. 32

Question 17 [Practice]

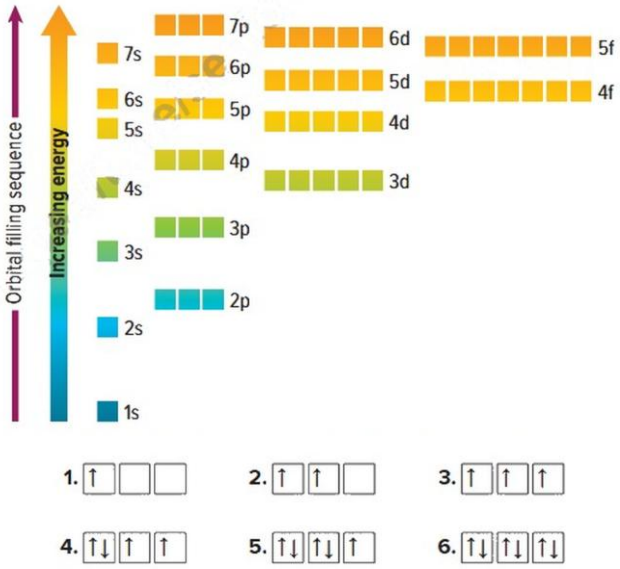
The principal quantum number (n) indicates the:

A. direction of electron spin
B. relative size and energy of orbitals
C. orientation of an orbital
D. shape of an orbital

Q3

Learning outcome: Explain the Aufbau principle, the Pauli exclusion principle, and Hund's rule

Textbook: pp. 23–24 – Figure 18 + Table 3



Key points from the lesson

- Aufbau principle: each electron occupies the lowest-energy orbital available: 1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p …
- Orbitals in the same sublevel have equal energy (the three 2p orbitals). The 4s sublevel is lower in energy than 3d, so it fills first.
- Pauli exclusion principle: an orbital holds at most two electrons, and they must have opposite spins [↑↓].
- Hund's rule: single electrons with the same spin occupy each equal-energy orbital before any orbital gets a second electron.
- Example (nitrogen $2p^3$): [↑][↑][↑] is correct; [↑↓][↑][ ] violates Hund's rule.

Question 1 [Past exam 2025–26]

Which electron configuration violates (does not comply with) the aufbau principle?

A. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$
B. $1s^2\ 2s^2\ 2p^6\ 3s^2$
C. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$
D. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Question 2 [Past exam 2025–26]

Which orbital diagram (2s | 2p) violates the Pauli exclusion principle?

A. [↑↑] [↑][↑][↑]
B. [↑↓] [↑↓][ ][ ]
C. [↑↓] [↑][↑][↑]
D. [↑] [↑][↑][↑]

Question 3 [Mock exam]

According to the aufbau principle, which sublevel fills with electrons before 3d?

A. 4s
B. 4p
C. 4d
D. 5s

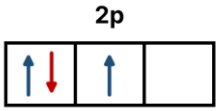
Question 4 [Mock exam]

Which diagram of the 2p electrons of nitrogen (N) follows Hund's rule and the Pauli principle?

A.



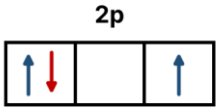
B.



C.



D.



Question 5 [Practice]

Which orbital diagram for $2p^3$ obeys Hund's rule?

A. [↑↓][↑][ ]
B. [↑][↑][↑]
C. [↑][↓][↑↓]
D. [↑↓][↑↓][↑↓]

Question 6 [Practice]

States that each electron occupies the lowest energy orbital available.

A. Pauli exclusion principle
B. Heisenberg uncertainty principle
C. Hund's rule
D. Aufbau principle

Question 7 [Practice]

"Electrons occupy orbitals of lowest energy first" is which rule?

A. Hund's rule
B. Aufbau principle
C. Pauli exclusion principle

Question 8 [Practice]

States that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbital.

A. Hund's rule
B. Aufbau principle
C. de Broglie equation
D. Pauli exclusion principle

Question 9 [Practice]

Which principle states that each electron occupies the lowest-energy orbital available?

A. Pauli exclusion principle
B. Heisenberg uncertainty principle
C. Hund's rule
D. Aufbau principle

Question 10 [Practice]

States that a maximum of two electrons can occupy a single atomic orbital but only if they have opposite spins.

A. Hund's rule
B. Pauli exclusion principle
C. Heisenberg uncertainty principle
D. Aufbau principle

Question 11 [Practice]

Which principle states that it is not possible to know precisely both the velocity and the position of a particle at the same time?

A. Hund's rule
B. Pauli exclusion principle
C. Heisenberg uncertainty principle
D. Aufbau principle

Question 12 [Practice]

According to the Pauli exclusion principle, how many electrons can occupy a single atomic orbital?

A. One
B. Two, with the same spin
C. Two, with opposite spins
D. Four

Question 13 [Practice]

Which rule states that electrons fill the lowest-energy orbitals first?

A. Hund's rule
B. Pauli exclusion principle
C. Aufbau principle
D. Octet rule

Question 14 [Practice]

The orbital [↑↑] (two electrons with the same spin) violates the:

A. Aufbau principle
B. Bohr model
C. Hund's rule
D. Pauli exclusion principle

Question 15 [Practice]

Which orbital diagram is correct for the $3p^3$ electrons of phosphorus?

A. [↑↓] [↑] [ ]
B. [↑] [↓] [↑]
C. [↑] [↑] [↑]
D. [↑↓] [↓] [ ]

Question 16 [Practice]

Using the aufbau diagram, which sublevel fills first?

A. 3d
B. 4p
C. 4d
D. 4s

Question 17 [Practice]

The diagram 1s[↑↓] 2s[↑↓] 2p[↑↓][↑↓][ ] (8 electrons) is wrong because it violates:

A. Aufbau principle
B. Pauli exclusion principle
C. Hund's rule
D. nothing – it is correct

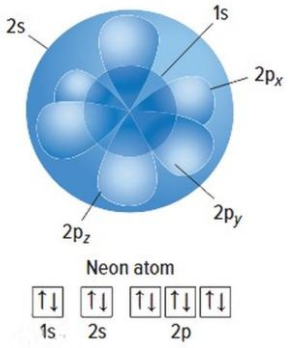
Q4

Learning outcome: Write the electron configuration, orbital diagram, and noble gas notation of different elements

Textbook: pp. 25–26 – Table 5

Table 5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or [Ar]



Key points from the lesson

- Electron configuration notation: sublevels with superscripts for the number of electrons: O = $1s^2 2s^2 2p^4$.
- Orbital diagram: one box per orbital, arrows for electrons; a full orbital is [↑↓].
- Noble-gas notation: symbol of the previous noble gas in brackets + the remaining electrons: Na = $[Ne]3s^1$, P = $[Ne]3s^2 3p^3$.
- The sum of the superscripts = atomic number (number of electrons).
- Exceptions (extra-stable half-filled / filled d): Cr = $[Ar]4s^1 3d^5$, Cu = $[Ar]4s^1 3d^{10}$.

Question 1 [Past exam 2025–26]

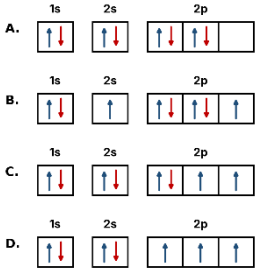
Using noble-gas notation, which represents the ground-state electron configuration of silicon (Si)?

A. $[Ne]3s^2 3p^2$
B. $[Ne]3s^2 3p^4$
C. $[Ar]4s^2 3d^2$
D. $[He]2s^2 2p^2$

Question 2 [Mock exam]

The correct electron configuration of sodium, Na (atomic number 11), is:

A. $1s^2\ 2s^2\ 2p^7$
B. $1s^2\ 2s^2\ 2p^6\ 3s^1$
C. $1s^2\ 2s^2\ 2p^6\ 3p^1$
D. $1s^2\ 2s^2\ 2p^5\ 3s^2$



Question 3 [Mock exam]

Which orbital diagram correctly represents an oxygen atom, O (atomic number 8)?

Question 4 [Practice]

Which orbital is filled after 2p?

A. 3s
B. 3p
C. 3d
D. 2d

Question 5 [Practice]

What must be true for two electrons to occupy the same orbital?

A. Same spin
B. Opposite spins
C. Different energy levels
D. Different sublevels

Question 6 [Practice]

Which orbital is filled after 4s?

A. 4p
B. 3d
C. 4d
D. 5s

Question 7 [Practice]

According to Hund's rule, electrons in orbitals of equal energy…

A. occupy each orbital singly before any orbital gets a second electron
B. pair up before moving to the next orbital
C. fill the lowest energy orbitals first
D. fill the highest energy orbitals first

Question 8 [Practice]

Which orbital is filled after 6s?

A. 6p
B. 5d
C. 4f
D. 7s

Question 9 [Practice]

In a set of p orbitals, how are electrons distributed according to Hund's rule?

A. One electron in each orbital with parallel spins before pairing
B. Two electrons in one orbital before the next
C. Randomly
D. Opposite spins first

Question 10 [Practice]

The orbital related to the 4s sublevel has…

A. lower energy than the 3d orbitals
B. higher energy than the 3d orbitals
C. the same energy as the 3d orbitals
D. no relation to 3d

Question 11 [Practice]

Light below the threshold frequency shines on a metal for a long time. What happens?

A. Electrons are ejected after a while
B. No photoelectrons are ejected
C. More electrons are ejected if the light is brighter
D. The metal emits photons of higher frequency

Question 12 [Practice]

Which symbol represents an orbital with one electron?

A. [↑↓]
B. [↑]
C. [ ]
D. [↓↓]

Question 13 [Practice]

What is the electron configuration of magnesium (Z = 12)?

A. $1s^2 2s^2 2p^6 3s^1$
B. $1s^2 2s^2 2p^6 3s^2$
C. $1s^2 2s^2 2p^8$
D. $[Ne]3p^2$

Question 14 [Practice]

Using Table 5, what is the noble-gas notation of silicon (Z = 14)?

Table 5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or [Ar]

A. $[Ne]3s^2 3p^2$
B. $[Ar]3s^2 3p^2$
C. $[He]2s^2 2p^2$
D. $[Ne]3s^2 3p^4$

Question 15 [Practice]

Which element has the configuration $[Ne]3s^2 3p^5$?

A. F
B. S
C. Cl
D. Ar

Question 16 [Practice]

How many unpaired electrons does a sulfur atom (…$3p^4$) have?

A. 0
B. 1
C. 2
D. 4

Question 17 [Practice]

What is the correct noble-gas notation for copper (Z = 29)?

A. $[Ar]4s^2 3d^9$
B. $[Ar]3d^{11}$
C. $[Kr]5s^1 4d^{10}$
D. $[Ar]4s^1 3d^{10}$

Q5

Learning outcome: Represent the electron dot structure of atoms

Textbook: pp. 28–29 – Table 6 + Example 3

Table 6 Electron Configurations and Dot Structures

Element	Atomic Number	Electron Configuration	Electron-Dot Structure
Lithium	3	$1s^2 2s^1$	Li·
Beryllium	4	$1s^2 2s^2$	·Be·
Boron	5	$1s^2 2s^2 2p^1$	·B·
Carbon	6	$1s^2 2s^2 2p^2$	·C·
Nitrogen	7	$1s^2 2s^2 2p^3$	·N·
Oxygen	8	$1s^2 2s^2 2p^4$	:O·
Fluorine	9	$1s^2 2s^2 2p^5$	:F·
Neon	10	$1s^2 2s^2 2p^6$	:Ne:

a. ·Se: b. ·Se· c. ·Se· d. ·S

Key points from the lesson

- Valence electrons are the electrons in the atom's outermost orbitals – the highest principal energy level.
- Electron-dot structure: the element's symbol surrounded by dots, one dot for each valence electron.
- Dots are placed one at a time on the four sides, then paired (N: 5 dots, O: 6 dots, Ne: 8 dots).
- Example 3: Se = $[Ar]4s^2 3d^{10} 4p^4 \rightarrow 4s^2 + 4p^4 = 6$ valence electrons; the 3d electrons are not valence electrons.

Question 1 [Past exam 2025–26]

Which of the following atoms contains 7 valence electrons? (Cl 17, Na 11, Mg 12, O 8)

A. Chlorine (Cl)
B. Sodium (Na)
C. Magnesium (Mg)
D. Oxygen (O)

Question 2 [Mock exam]

How many dots are in the electron-dot structure of chlorine, Cl (atomic number 17)?

A. 1
B. 5
C. 17
D. 7

Question 3 [Practice]

What is an electron dot structure?

A. A diagram that shows the symbol of an element surrounded by dots representing its valence electrons
B. A diagram that shows all the electrons of an atom as dots
C. A diagram showing the protons and neutrons in the nucleus
D. A list of an atom's energy levels and sublevels

Question 4 [Practice]

A solid found in emeralds with the dot structure ·X· is… (Practice 26)

A. carbon
B. sulfur
C. beryllium
D. cesium

Question 5 [Practice]

What does each dot in the electron cloud represent?

A. A location where the electron could be at a given moment
B. The exact position of an electron
C. The energy of an electron
D. The velocity of an electron

Question 6 [Practice]

A gas from the list H, He, N, O, F, Cl, Ne with the dot structure ·X· is…

A. hydrogen
B. helium
C. nitrogen
D. neon

Question 7 [Practice]

Which atom is usually the central atom in a Lewis structure?

A. The atom with the least attraction for shared electrons
B. The atom with the most electrons
C. Always hydrogen
D. The atom closest to fluorine

Question 8 [Practice]

Consists of an element's symbol (nucleus and inner electrons) surrounded by dots representing the atom's valence electrons.

A. energy sublevel
B. photon
C. quantum
D. electron-dot structure

Question 9 [Practice]

What does each dot in an electron-cloud diagram represent?

A. A location where the electron could be at a given moment
B. The exact position of an electron
C. The energy of an electron
D. The velocity of an electron

Question 10 [Practice]

How many hydrogen atoms are in a molecule of phosphine, PH_3?

A. 1
B. 2
C. 3
D. 4

Question 11 [Practice]

The molecular formula of a compound tells you…

A. the type and number of each atom in a molecule
B. the shape of the molecule
C. where the lone pairs are
D. the bond lengths

Question 12 [Practice]

How many dots are in the electron-dot structure of carbon?

A. 2
B. 4
C. 6
D. 12

Question 13 [Practice]

Selenium is $[Ar]4s^23d^{10}4p^4$. How many dots appear in its electron-dot structure?

A. 4
B. 16
C. 6
D. 34

Question 14 [Practice]

An atom has 13 electrons. Which electron-dot structure represents it?

A. Al with 3 dots
B. C with 4 dots
C. Al with 13 dots
D. Mg with 2 dots

Question 15 [Practice]

Which element has an electron-dot structure with 7 dots?

A. Ne
B. N
C. O
D. F

Question 16 [Practice]

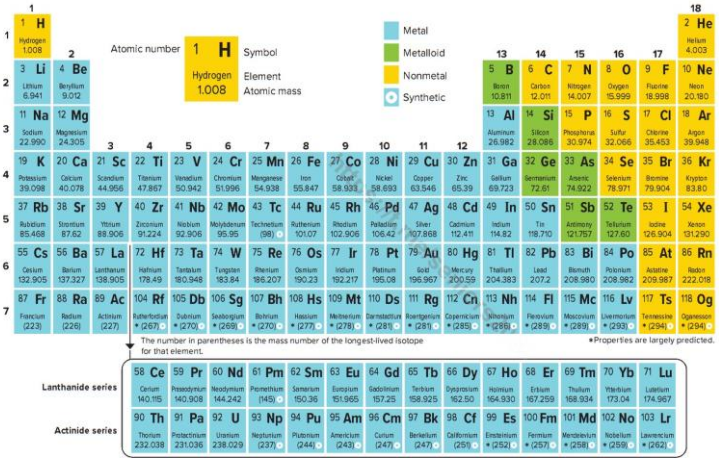
How many valence electrons are in $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$?

A. 36
B. 18
C. 10
D. 8

Q6

Learning outcome: List the families of the periodic table with the group number and the elements

Textbook: pp. 40–41



Key points from the lesson

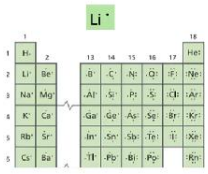
- Columns are groups (1–18); rows are periods. Elements in a group have similar chemical properties.
- Group 1 (except H): alkali metals; group 2: alkaline earth metals (Ca, Mg) – both highly reactive.
- Groups 3–12: transition metals; lanthanide and actinide series: inner transition metals.
- Group 17: halogens (very reactive, e.g., F in toothpaste); group 18: noble gases (unreactive, used in lasers and neon signs).
- Metals on the left (blue), nonmetals upper right (yellow; Br is the only liquid nonmetal), metalloids (green) along the staircase, e.g., Si, Ge.

Question 1 [Past exam 2025–26]

Lithium (Li) has the electron-dot structure Li·. Using the periodic table, what is an expected property of this element?

Lithium element (Li) has the following Electron-dot structure.
Using the periodic table below, what are the expected chemical and physical properties for this element?

عنصر الليثيوم (Li) له الترميز النقطي التالي.
باستخدام الجدول الدوري أدناه، ما الخواص الكيميائية والفيزيائية المتوقعة لهذا العنصر؟



A. Highly chemically reactive
B. Strong and lightweight
C. Brittle and dull
D. Poor conductor of heat and electricity

Question 2 [Mock exam]

The vertical columns in the modern periodic table are called:

A. Periods
B. Orbitals
C. Groups
D. Metals

Question 3 [Mock exam]

Which is a general property of the alkali metals (group 1)?

A. Very reactive and react with water
B. Unreactive gases
C. Have 7 valence electrons
D. Brittle nonmetals

Question 4 [Mock exam]

Which element is a halogen and is used to disinfect drinking water?

A. Sodium Na
B. Neon Ne
C. Magnesium Mg
D. Chlorine Cl

Question 5 [Mock exam]

Elements in the same group have similar chemical properties because they have the same:

A. Atomic number
B. Number of valence electrons
C. Atomic mass
D. Number of energy levels

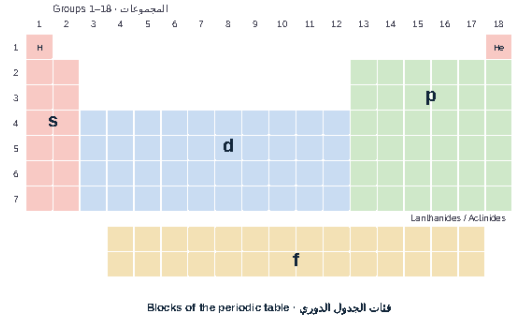
Question 6 [Mock exam]

Which is a general property of nonmetals?

A. Shiny
B. Malleable and ductile
C. Poor conductors of heat and electricity
D. Good electrical conductors

Question 7 [Practice]

Which periodic table group has 2 electrons in its outermost energy level?



A. group 1 (alkali metals)
B. group 17 (halogens)
C. group 16 (oxygen family)
D. group 2 (alkaline earth metals)

Question 8 [Practice]

Which of the following elements has the lowest electronegativity value?

A. Cesium
B. Sodium
C. Potassium
D. Francium

Question 9 [Practice]

An element has the noble-gas notation [Ar]$4s^2 3d^6$. It is…

A. a representative element in group 6
B. a noble gas
C. a transition element in period 4
D. an inner transition element

Question 10 [Practice]

Each family in the periodic table has its own characteristic properties based on the number of:

Use the figure above

A. neutrons
B. valence electrons
C. protons
D. ions

Question 11 [Practice]

The element boron is directly above aluminum in the periodic table. Which statement about boron is true?

Use the figure above

A. Boron is in the same period as aluminum.
B. Boron is in the same group as aluminum and has 5 valence electrons.
C. Boron has 5 valence electrons and is in the same period as aluminum.
D. Boron is in the same group as aluminum and has 3 valence electrons.

Question 12 [Practice]

Which statement correctly compares representative and transition elements?

A. Representative elements are all in the d-block
B. Transition elements include the noble gases
C. Both groups contain only metals
D. Representative elements include metals, nonmetals and metalloids; transition elements are all metals

Question 13 [Practice]

Each family in the periodic table has its own characteristic properties based on its number of valence electrons.

Use the figure above

A. True
B. False

Question 14 [Practice]

Without looking at a periodic table, what is the expected outer electron configuration of the element in group 7, period 4?

Use the figure above

A. $4s^2 3d^5$
B. $4s^2 4p^5$
C. $4s^2 3d^7$
D. $4s^1 3d^6$

Question 15 [Practice]

The oxidation number of an element in an ionic compound equals…

A. the number of electrons transferred from the atom to form the ion
B. the number of protons
C. the group number always
D. the number of energy levels

Question 16 [Practice]

Group 17 elements are called:

A. noble gases
B. alkaline earth metals
C. alkali metals
D. halogens

Question 17 [Practice]

Calcium and magnesium belong to the:

A. alkali metals (group 1)
B. transition metals
C. alkaline earth metals (group 2)
D. halogens

Question 18 [Practice]

The lanthanide and actinide series are known as:

A. inner transition metals
B. metalloids
C. noble gases
D. alkali metals

Question 19 [Practice]

Silicon, used in computer chips, is classified as a:

A. nonmetal
B. metalloid
C. noble gas
D. transition metal

Question 20 [Practice]

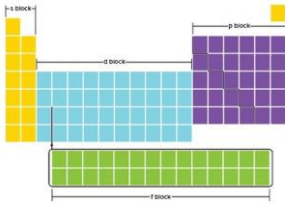
Group 18 elements are used in lasers and neon signs because they are:

A. brittle solids
B. very reactive metals
C. good conductors
D. extremely unreactive

Q7

Learning outcome: Differentiate between representative elements and transition elements

Textbook: p. 40

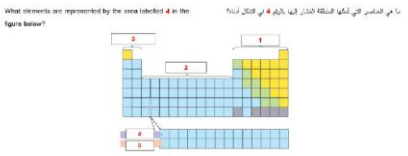


Key points from the lesson

- Representative elements: groups 1, 2 and 13–18 (s-block and p-block). They include metals, nonmetals and metalloids.
- They show a wide range of physical and chemical properties.
- Transition elements: groups 3–12, divided into transition metals (d-block) and inner transition metals (f-block: lanthanides, actinides).
- All transition elements are metals.
- Valence electrons of representative elements: groups 1–2 = group number; groups 13–18 = group number − 10 (He has 2).

Question 1 [Past exam 2025–26]

What elements are represented by the area labelled 4 in the figure?



A. Lanthanide series
B. Alkaline earth metals
C. Actinide series
D. Halogens

Question 2 [Mock exam]

In the periodic table, the transition elements are in groups:

A. 1 and 2
B. 3 to 12
C. 13 to 17
D. 18 only

Question 3 [Practice]

As an electron moves from n = 1 to n = 4, energy is...

A. absorbed
B. released

Question 4 [Practice]

Iron (Fe) is classified as…

A. a transition metal
B. a metalloid
C. a representative element
D. a noble gas

Question 5 [Practice]

As an electron moves from n = 4 to n = 1, energy is...

A. absorbed
B. released

Question 6 [Practice]

Which element is a representative element?

A. Zinc (Zn)
B. Calcium (Ca)
C. Copper (Cu)
D. Chromium (Cr)

Question 7 [Practice]

Electrons can exist in between energy levels.

A. True B. False

Question 8 [Practice]

The inner transition metals (lanthanides and actinides) belong to the…

A. p-block
B. s-block
C. d-block
D. f-block

Question 9 [Practice]

Which electron jump gives the more energetic light?

A. 5th to the 2nd energy level B. 3rd to the 2nd energy level

Question 10 [Practice]

A representative element has three valence electrons. In which group is it located?

A. Group 1 B. Group 13 C. Group 3 D. Group 15

Question 11 [Practice]

Which strong, light transition metal is used to make frames of bicycles and eyeglasses?

A. Iron (Fe)
B. Titanium (Ti)
C. Copper (Cu)
D. Zinc (Zn)

Question 12 [Practice]

Which groups contain the representative elements?

A. Groups 1, 2 and 13–18
B. Lanthanides and actinides
C. Group 18 only
D. Groups 3–12

Question 13 [Practice]

The transition elements are divided into:

A. alkali and alkaline earth metals
B. metals and nonmetals
C. transition metals and inner transition metals
D. halogens and noble gases

Question 14 [Practice]

How many valence electrons does an element in group 15 have?

A. 3 B. 5 C. 10 D. 15

Question 15 [Practice]

Which statement is correct?

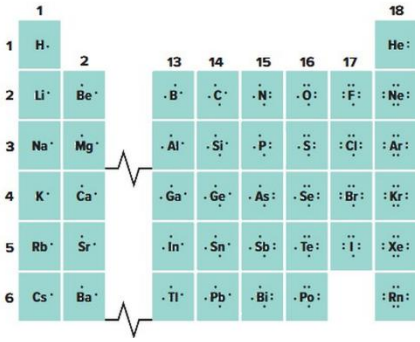
A. Transition elements are in the p-block.
B. All transition elements are metals.
C. Halogens are transition elements.
D. All representative elements are nonmetals.

Q8

Learning outcome: Use the electron configuration, orbital and noble gas notation of an element (Z 1–36) to locate it (period, group, block)

Textbook: pp. 44–48 – Example 1 + Figure 9





Key points from the lesson

- Period = the energy level (n) of the valence electrons. Example: Ga $[Ar]4s^2 3d^{10} 4p^1$ → period 4.
- Block = the sublevel being filled last (s, p, d or f).
- Group: s-block → number of valence electrons (ns^1 = 1, ns^2 = 2); p-block → 10 + number of valence electrons (ns^2np^4 → 16).
- d-block: group = ns electrons + (n−1)d electrons. Example: $[Ar]4s^2 3d^6$ → group 8.
- Example 1: Sr $[Kr]5s^2$ → period 5, group 2, s-block.

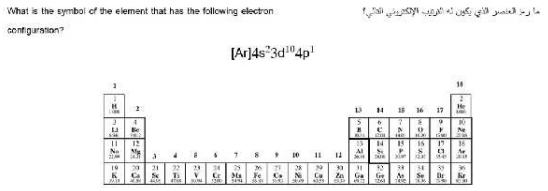
Question 1 [Past exam 2025–26]

To which period, group and block does the element $[Ar]4s^2 3d^{10} 4p^4$ belong?

A. Period 4 · group 16 · p-block
B. Period 5 · group 12 · d-block
C. Period 4 · group 6 · p-block
D. Period 5 · group 16 · d-block

Question 2 [Past exam 2025–26]

What is the symbol of the element with the electron configuration $[Ar]4s^2 3d^{10} 4p^1$?



A. Ga
B. Al
C. Si
D. Ge

Question 3 [Past exam 2025–26]

An element is in group 17, period 4. What is its electron configuration?

A. $[Ar]4s^2 3d^{10} 4p^5$
B. $[Ar]4s^2 3d^{10} 4p^6$
C. $[Ar]4s^2 3d^{10} 4p^3$
D. $[Ar]4s^2 3d^5$

Question 4 [Mock exam]

An element has the configuration [Ne] $3s^2$ $3p^4$. How many valence electrons does it have?

A. 6
B. 2
C. 4
D. 16

Question 5 [Mock exam]

An element has the configuration [Ar] $4s^2$. Where is it in the periodic table?

A. Period 2, group 4
B. Period 3, group 2
C. Period 4, group 12
D. Period 4, group 2

Question 6 [Practice]

What is the noble-gas shorthand for a sulfur atom?

A. [Ar] $3p^4$
B. [He] $3s^2$ $3p^4$
C. [Ne] $3s^2$ $3p^4$
D. [Na] $3s^2$ $3p^4$

Question 7 [Practice]

An element has the electron configuration [Ne]$3s^2 3p^4$. Give its group, period and block.

A. Group 16, period 3, p-block
B. Group 6, period 3, p-block
C. Group 16, period 4, p-block
D. Group 14, period 3, s-block

Question 8 [Practice]

An element has the electron configuration [Xe]$6s^2 4f^{14} 5d^{10} 6p^5$. Give its group, period and block.

A. Group 17, period 6, f-block
B. Group 7, period 6, p-block
C. Group 17, period 6, p-block
D. Group 15, period 6, p-block

Question 9 [Practice]

Which statement is true about electronegativity?

A. It decreases across a period from left to right.
B. It is a measure of an atom's ability to attract electrons in a chemical bond.
C. It is the same for all elements.
D. It is measured in kilojoules per mole (kJ/mol).

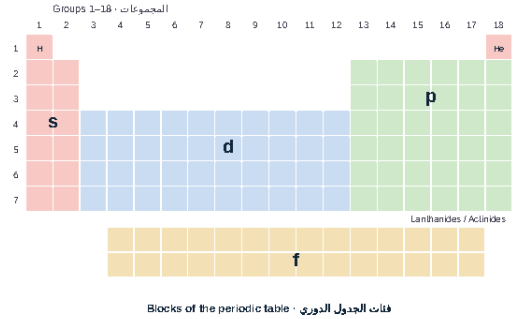
Question 10 [Practice]

An element has the electron configuration [Kr]$5s^2 4d^1$. Give its group, period and block.

A. Group 2, period 5, s-block
B. Group 3, period 5, d-block
C. Group 3, period 4, d-block
D. Group 13, period 5, p-block

Question 11 [Practice]

In the periodic table, the vertical columns are called:



A. groups
B. periods
C. rows
D. blocks

Question 12 [Practice]

An element has the electron configuration $[Ar]4s^2 3d^{10} 4p^3$. Give its group, period and block.

A. Group 5, period 4, d-block
B. Group 15, period 3, p-block
C. Group 15, period 4, p-block
D. Group 13, period 4, p-block

Question 13 [Practice]

What are the period, group and block of $[Ne]3s^2$?

A. Period 2, group 2, s-block
B. Period 3, group 2, s-block
C. Period 3, group 12, d-block
D. Period 2, group 3, p-block

Question 14 [Practice]

Where is the element $[Ar]4s^2 3d^{10} 4p^5$ located?

A. Period 4, group 7, d-block
B. Period 3, group 17, p-block
C. Period 4, group 17, p-block
D. Period 5, group 15, p-block

Question 15 [Practice]

Which configuration belongs to an element in group 4, period 4, d-block?

A. $[Ne]3s^2 3p^2$
B. $[Ar]4s^2 3d^2$
C. $[Ar]4s^2 3d^5$
D. $[Ar]4s^2 3d^{10} 4p^2$

Question 16 [Practice]

$[Ne]3s^2 3p^4$ is located in:

A. group 16, period 3, s-block
B. group 3, period 6, p-block
C. group 4, period 3, p-block
D. group 16, period 3, p-block

Question 17 [Practice]

An element is in period 4, group 1. Its configuration is:

A. $[Ar]4s^1$
B. $[Kr]5s^1$
C. $[Ar]4s^2$
D. $[Ne]3s^1$

Q9

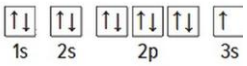
Learning outcome: The electron configuration: filling of sublevels

Textbook: p. 46

Table 4 Noble Gas Electron Configuration

Period	Principal Energy Level	Element	Electron Configuration
1	$n = 1$	helium	$1s^2$
2	$n = 2$	neon	$[He]2s^2p^6$
3	$n = 3$	argon	$[Ne]3s^2 3p^6$
4	$n = 4$	krypton	$[Ar]4s^2 3d^{10} 4p^6$

$1s^2 2s^2 2p^6 3s^1$



Key points from the lesson

- s-block: groups 1–2 and helium; an s sublevel holds 2 electrons, so the s-block spans 2 groups.
- p-block: groups 13–18; three p orbitals hold 6 electrons → 6 groups. No p-block in period 1 (no 1p sublevel).
- d-block: filled ns orbital and filling (n−1)d orbitals. 4s fills before 3d: Sc = $[Ar]4s^2 3d^1$, Ti = $[Ar]4s^2 3d^2$.
- Five d orbitals hold 10 electrons → the d-block spans 10 groups.
- f-block: inner transition metals; seven f orbitals hold 14 electrons → 14 columns.

Question 1 [Practice]

Ground-state configuration of bromine (Practice 19a):

A. $[Ar]\ 4s^2\ 3d^{10}\ 4p^5$
B. $[Ar]\ 4s^2\ 4p^5$
C. $[Kr]\ 4p^5$
D. $[Ar]\ 3d^{10}\ 4p^7$

Question 2 [Practice]

Which sublevels of the 3rd energy level are filled in krypton? (Booklet Q106b)

A. 3s only
B. 3s and 3p
C. 3s, 3p and 3d
D. none

Question 3 [Practice]

Why is Cd^{2+} stable even though it does not have a noble-gas configuration?

A. It gained two electrons
B. It has a half-filled d sublevel
C. It has 8 electrons in its d sublevel
D. Its outer s, p and d sublevels are completely filled (18 electrons)

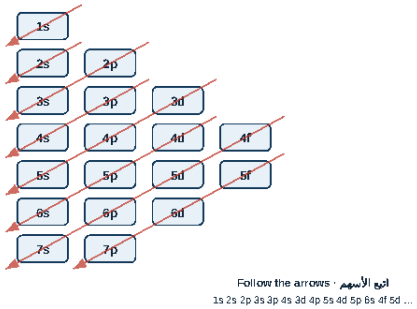
Question 4 [Practice]

Ground-state configuration of strontium (Practice 19b):

A. $[Ar]\ 5s^2$
B. $[Kr]\ 5s^2$
C. $[Kr]\ 4d^2$
D. $[Xe]\ 5s^2$

Question 5 [Practice]

According to the aufbau diagram, which sublevel is filled after 2p?



A. 2d
B. 3s
C. 3p
D. 2s

Question 6 [Practice]

According to the aufbau diagram, which sublevel is filled after 6s?

Use the figure above

A. 6p
B. 5d
C. 4f
D. 7s

Question 7 [Practice]

Ground-state configuration of titanium (Practice 19f):

A. [Ar] $4s^2$ $3d^2$
B. [Ar] $3d^4$
C. [Ar] $4s^2$ $4p^2$
D. [Ne] $4s^2$ $3d^2$

Question 8 [Practice]

According to the aufbau diagram, which sublevel is filled after 4s?

Use the figure above

A. 4p
B. 3d
C. 4d
D. 5s

Question 9 [Practice]

Which sublevels of the 3rd energy level are filled in the element krypton (Z = 36)?

A. 3s and 3p only
B. 3s, 3p and 3d
C. 3s only
D. 3d only

Question 10 [Practice]

The correct ground-state configuration of chromium is…

A. [Ar] $4s^2$ $3d^4$
B. [Ar] $4s^1$ $3d^5$
C. [Ar] $3d^6$
D. [Ar] $4s^2$ $4p^4$

Question 11 [Practice]

How many columns wide is the f-block?

A. 2
B. 6
C. 10
D. 14

Question 12 [Practice]

Why does the p-block span six groups?

A. It contains six periods.
B. There are six p orbitals in each level.
C. p-block elements have six protons.
D. The three p orbitals hold a maximum of six electrons.

Question 13 [Practice]

What is the electron configuration of titanium (Z = 22)?

A. $[Ar]4s^1 3d^3$
B. $[Ar]4s^2 4p^2$
C. $[Ar]4s^2 3d^2$
D. $[Ar]3d^4$

Question 14 [Practice]

Why are there no p-block elements in period 1?

A. Period 1 is in the d-block.
B. The p sublevel does not exist for n = 1.
C. 1p is too high in energy.
D. H and He are metals.

Question 15 [Practice]

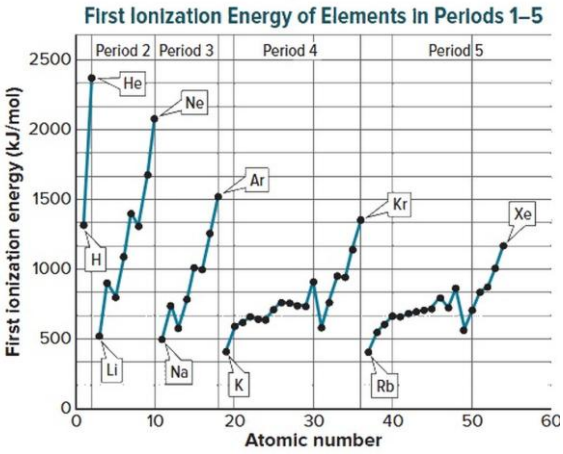
d-block elements are characterized by:

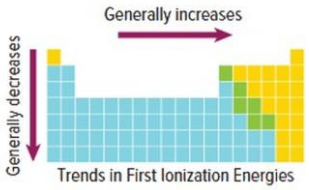
A. partly filled p orbitals
B. filled f orbitals
C. a filled ns orbital and filling (n−1)d orbitals
D. only one s electron

Q10

Learning outcome: Explain the trend of first ionization energy across a period (exceptions between groups 2 & 3 and 5 & 6) and down a group

Textbook: p. 54 – Figure 17





Key points from the lesson

- Ionization energy: the energy required to remove an electron from a gaseous atom. The first ionization energy removes the first outermost electron.
- Across a period, first ionization energy generally increases: increasing nuclear charge holds valence electrons more tightly.
- Down a group, it decreases: atomic size increases, so valence electrons are farther from the nucleus.
- Exceptions: group 13 < group 2 (B 800 < Be 900 kJ/mol: 2p electron is easier to remove than 2s); group 16 < group 15 (O 1310 < N 1400: paired 2p electrons repel).
- Noble gases have the highest values (peaks in Figure 17); alkali metals the lowest (e.g., Li in batteries).

Question 1 [Past exam 2025–26]

Which describes the trends in atomic radii correctly?

A. Across a period: generally decrease · Down a group: generally increase
B. Across a period: generally increase · Down a group: generally increase
C. Across a period: generally decrease · Down a group: generally decrease
D. Across a period: generally increase · Down a group: generally decrease

Question 2 [Past exam 2025–26]

Which of the following explains the general decrease in first ionization energy as you move down a group?

A. Atomic size increases as you move down the group
B. The nuclear charge decreases drastically
C. More energy is required to remove valence electrons farther from the nucleus
D. Atoms lose protons as they move down the group

Question 3 [Mock exam]

Which element has the largest atomic radius?

A. Na
B. K
C. Mg
D. Cl

Question 4 [Mock exam]

The first ionization energy is the energy needed to:

A. Remove an electron from a gaseous atom
B. Add an electron to an atom
C. Break a bond between two atoms
D. Melt a solid

Question 5 [Mock exam]

Which statement about first ionization energy is correct?

A. Na has a higher IE than Li
B. O has a higher IE than N
C. Be has a higher IE than B
D. F has a higher IE than Ne

Question 6 [Practice]

Which element would have the highest ionization energy in Period 3?

A. Sodium (Na)
B. Magnesium (Mg)
C. Aluminum (Al)
D. Argon (Ar)

Question 7 [Practice]

As you move from left to right across a period, the ionization energy:

A. Decreases
B. Fluctuates
C. Remains the same
D. Increases

Question 8 [Practice]

A new element has ionization energies IE1 127, IE2 287, IE3 625, IE4 2651, IE5 3540 kJ/mol. Which group does it belong to?

A. Group 1
B. Group 2
C. Group 13
D. Group 15

Question 9 [Practice]

As you move down a group, the ionization energy:

A. Decreases
B. Increases
C. Remains the same
D. Fluctuates

Question 10 [Practice]

Electronegativity is measured on a relative scale. What is the highest value on this scale?

A. 3.98
B. 4.00
C. 2.55
D. 1.00

Question 11 [Practice]

Which element would have the lowest ionization energy in Group 1?

A. Lithium (Li)
B. Sodium (Na)
C. Potassium (K)
D. Cesium (Cs)

Question 12 [Practice]

Why does ionization energy increase across a period from left to right?

A. The number of protons in the nucleus increases
B. The number of protons in the nucleus decreases
C. The atomic radius increases
D. The number of neutrons in the nucleus increases

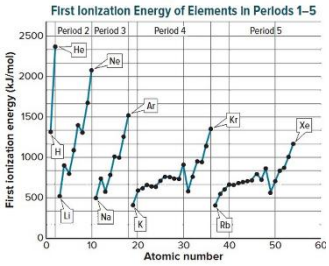
Question 13 [Practice]

Which period-3 element has the highest first ionization energy?

A. Na
B. Al
C. P
D. Cl

Question 14 [Practice]

Using Figure 17, which elements have the highest first ionization energy in each period?



A. transition metals
B. alkali metals
C. noble gases
D. halogens

Question 15 [Practice]

Why does first ionization energy decrease down a group?

A. Nuclear charge decreases.
B. Valence electrons decrease.
C. Atomic size increases; valence electrons are farther from the nucleus.
D. Atoms lose protons.

Question 16 [Practice]

Why is the first ionization energy of boron (800) lower than that of beryllium (900 kJ/mol)?

A. B is a metal.
B. Be has a larger radius.
C. B has fewer protons.
D. B's outer electron is in 2p, which is higher in energy than 2s.

Question 17 [Practice]

Which element has the lower first ionization energy, N or O, and why?

A. N, because it has fewer electrons
B. They are equal
C. N, because it is smaller
D. O, because of repulsion between its paired 2p electrons

Question 18 [Practice]

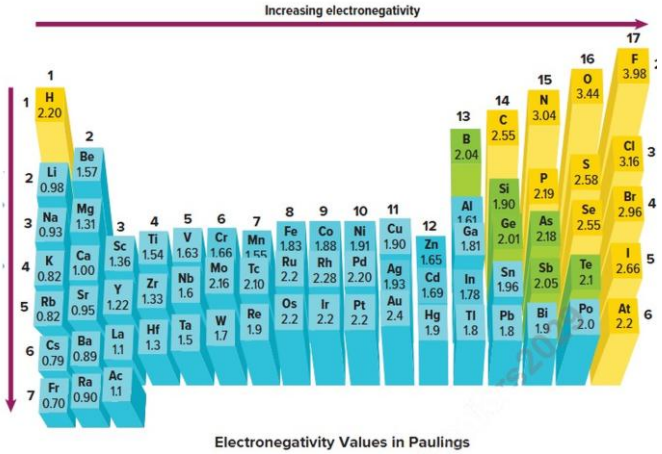
Lithium is used in batteries that deliver power quickly. Which property explains this?

A. high ionization energy
B. high electronegativity
C. it is a noble gas
D. low ionization energy – it loses electrons easily

Q11

Learning outcome: Predict the periodic trend of electronegativity across a period and down a group

Textbook: pp. 56–57 – Figure 19

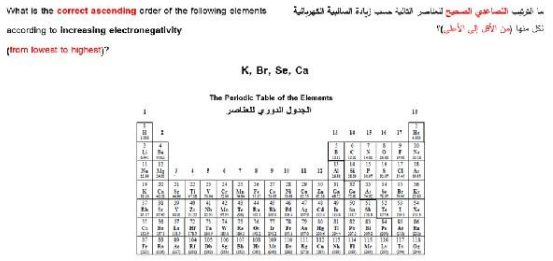


Key points from the lesson

- Electronegativity: the relative ability of an atom to attract electrons in a chemical bond (units: Paulings).
- It generally increases from left to right across a period (nuclear charge increases).
- It generally decreases down a group (atoms become larger).
- Fluorine is the most electronegative element (3.98); cesium (0.79) and francium (0.70) are the least.
- Noble gases have no values because they form very few compounds. In a bond, the more electronegative atom attracts the electrons more strongly.

Question 1 [Past exam 2025–26]

What is the correct order of K, Br, Se and Ca by increasing electronegativity (lowest to highest)?



A. K < Ca < Se < Br
B. Se < Br < Ca < K
C. Ca < K < Br < Se
D. Br < Se < Ca < K

Question 2 [Mock exam]

Which element has the highest electronegativity?

A. Na
B. O
C. Cl
D. F

Question 3 [Practice]

As you move from left to right across a period, electronegativity generally:

A. decreases
B. remains the same
C. increases
D. fluctuates randomly

Question 4 [Practice]

Which one of the following correctly describes the trend in electronegativity?

A. It increases across a period and decreases down a group.
B. It decreases across a period and decreases down a group.
C. It increases across a period and increases down a group.
D. It decreases across a period and increases down a group.

Question 5 [Practice]

Which of the following statements about electronegativity is true?

A. It decreases across a period from left to right.
B. It is a measure of an atom's ability to attract electrons in a chemical bond.
C. It is the same for all elements.
D. It is measured in kilojoules per mole (kJ/mol).

Question 6 [Practice]

How does electronegativity change across a period and down a group?

A. It increases across a period and decreases down a group.
B. It decreases across a period and increases down a group.
C. It increases across a period and down a group.
D. It does not change across a period.

Question 7 [Practice]

As you move down a group, electronegativity generally:

A. decreases
B. remains the same
C. increases
D. fluctuates randomly

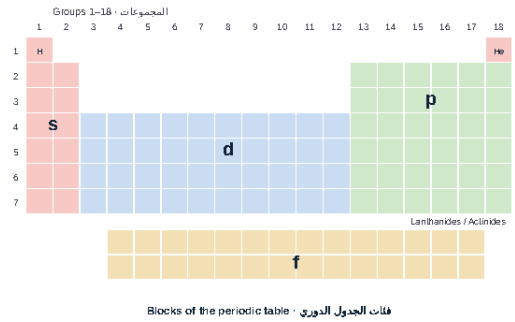
Question 8 [Practice]

Across a period, how does the size of the positive ions change, and what happens beginning with group 15 or 16?

A. Positive ion size decreases gradually; then the much larger negative ions begin and also decrease in size.
B. Positive ion size increases; then the negative ions begin and decrease in size.
C. Positive ion size stays constant; negative ions increase in size.
D. All ions increase in size across the period.

Question 9 [Practice]

As you move down a group in the periodic table, what happens to the atomic radius?



Blocks of the periodic table · فئات الجدول الدوري

A. It decreases.
B. It increases.
C. It remains the same.
D. It first increases, then decreases.

Question 10 [Practice]

What is the electronegativity value assigned to fluorine?

A. 2.98
B. 3.44
C. 3.98
D. 4.00

Question 11 [Practice]

From left to right across a period, how do atomic radius, first ionization energy and electronegativity change?

A. Radius decreases; ionization energy increases; electronegativity increases
B. Radius increases; ionization energy decreases; electronegativity decreases
C. Radius decreases; ionization energy decreases; electronegativity increases
D. Radius increases; ionization energy increases; electronegativity increases

Question 12 [Practice]

Which element is the most electronegative?

A. F
B. O
C. Cl
D. N

Question 13 [Practice]

Down a group, electronegativity generally:

A. stays constant
B. decreases
C. increases
D. doubles

Question 14 [Practice]

Te, O, Se and S are in group 16. Which order shows increasing electronegativity?

A. O < S < Se < Te
B. Te < Se < S < O
C. S < O < Te < Se
D. Se < Te < O < S

Question 15 [Practice]

Why are no electronegativity values given for the noble gases?

A. They are very reactive.
B. Their atoms are very large.
C. They are metals.
D. They form very few compounds.

Question 16 [Practice]

In an H–Cl bond, which atom attracts the shared electrons more strongly?

A. H, because it is smaller
B. Cl, because 3.16 > 2.20
C. Both equally
D. Neither

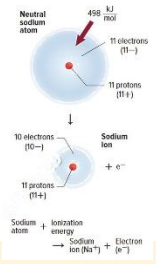
Q12

Learning outcome: Describe how ions (cations and anions) form to fulfill the octet rule

Textbook: p. 65

Table 2 Group 1, 2, and 13 Ions

Group	Configuration	Charge of Ion Formed
1	[noble gas]ns^1	1+ when the s^1 electron is lost
2	[noble gas]ns^2	2+ when the s^2 electrons are lost
13	[noble gas]ns^2np^1	3+ when the s^2p^1 electrons are lost



Key points from the lesson

- Octet rule: atoms gain, lose or share electrons to acquire a full set of eight valence electrons (noble-gas configuration).
- Metals lose valence electrons to form positive ions (cations). $Na \rightarrow Na^+ + e^-$ requires energy (ionization energy, 498 kJ/mol).
- Na^+ (11 protons, 10 electrons) has the configuration of neon, but it is still sodium – a +1 ion.
- Group 1 → 1+, group 2 → 2+, group 13 → 3+ (Table 2).
- Transition metals (outer ns^2) commonly form 2+ or 3+ ions: Fe^{2+} and Fe^{3+} (Fe^{3+}: loses two 4s and one 3d electron).

Question 1 [Past exam 2025–26]

Which correctly explains why sodium forms a cation (Na^+)? $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

A. Sodium loses one electron to attain a stable configuration like neon
B. Sodium gains one electron to achieve an octet
C. Sodium shares one electron with chlorine to achieve an octet
D. Sodium transfers two electrons to chlorine

Question 2 [Mock exam]

The octet rule states that atoms lose, gain, or share electrons to get:

A. 8 protons in the nucleus
B. 8 electrons in the outer energy level
C. Only 2 electrons in all levels
D. 18 electrons in the outer level

Question 3 [Mock exam]

When a magnesium atom, Mg, loses two electrons, it forms:

A. Positive ion Mg^{2+}
B. Negative ion Mg^{2-}
C. Positive ion Mg^+
D. Neutral atom Mg

Question 4 [Practice]

A pseudo-noble gas configuration is an outer arrangement of…

A. 18 electrons: $ns^2\ np^6\ nd^{10}$
B. 8 electrons: $ns^2\ np^6$
C. 2 electrons: ns^2
D. an empty outer level

Question 5 [Practice]

Gallium (Ga, $[Ar]4s^23d^{10}4p^1$) forms Ga^{3+}. Its configuration is…

A. [Ar] — a true noble-gas configuration
B. [Kr]
C. $[Ar]3d^{10}$ — a pseudo-noble gas configuration
D. $[Ar]4s^23d^8$

Question 6 [Practice]

When a transition metal atom forms a positive ion, which electrons are usually removed first?

A. the outer s electrons
B. the 1s electrons
C. the inner p electrons
D. the d electrons

Question 7 [Practice]

Which ion has a pseudo-noble gas configuration?

A. Cl^-
B. Mg^{2+}
C. Ag^+
D. Na^+

Question 8 [Practice]

When magnesium forms an ion, it:

A. loses two electrons → Mg^{2+}
B. shares two electrons
C. gains two electrons → Mg^{2-}
D. loses one electron → Mg^+

Question 9 [Practice]

Na^+ has the same electron configuration as:

A. Mg
B. Ne
C. Ar
D. K

Question 10 [Practice]

Which statement about cations is correct?

A. They form when atoms gain electrons.
B. They have more electrons than protons.
C. They have more protons than electrons.
D. They are negatively charged.

Question 11 [Practice]

Which is correct about forming a positive ion from a neutral atom?

A. The atom gains protons.
B. The atom gains electrons and energy is released.
C. The atom loses valence electrons and this requires energy.
D. The atom loses electrons and energy is released.

Question 12 [Practice]

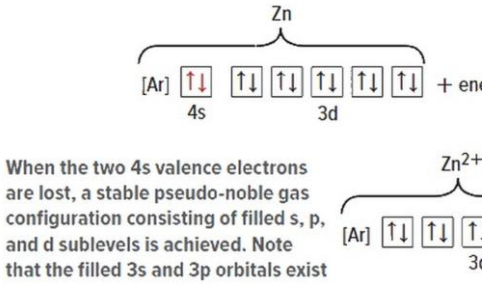
Iron ($[Ar]4s^23d^6$) forms Fe^{3+} by:

A. losing one 3d electron only
B. losing three 4s electrons
C. gaining three electrons
D. losing two 4s electrons and one 3d electron

Q13

Learning outcome: Describe how ions form to fulfill the octet rule – pseudo noble gas configuration

Textbook: p. 66 – Figure 3



Key points from the lesson

- An octet is the most stable configuration, but other configurations also provide some stability.
- Elements in groups 11–14 lose electrons to form an outer energy level with full s, p and d sublevels: a pseudo-noble-gas configuration.
- Zn = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$ loses its two 4s electrons → Zn^{2+} = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$.
- Electrons are removed from the highest energy level (4s) before 3d. ZnI_2 contains Zn^{2+} ions with this configuration.
- Other examples: Cd^{2+} = $[Kr]4d^{10}$, Cu^{+} = $[Ar]3d^{10}$.

Question 1 [Mock exam]

The noble-gas notation for the sodium ion, Na^+, is:

A. [Ne] $3s^1$
B. [Ne]
C. [Ne] $3s^2$
D. [Ar]

Question 2 [Practice]

Why do atoms form chemical bonds according to the octet rule?

A. To lose electrons and become positively charged.
B. To gain electrons and become negatively charged.
C. To complete their outer shell and attain a noble gas configuration.
D. To increase their atomic number.

Question 3 [Practice]

Booklet Q58–59 (carbonyl group): structure I has 4 lone-pair electrons on O and structure II has 6. Is the number of lone-pair electrons the same?

A. Yes B. No

Question 4 [Practice]

SF_6 is an example of…

A. an odd number of electrons
B. a noble gas
C. an expanded octet
D. a suboctet

Question 5 [Practice]

A suboctet is…

A. an ion with a negative charge
B. a molecule with resonance
C. a stable configuration with fewer than eight electrons around an atom
D. an atom with more than eight electrons

Question 6 [Practice]

Atoms of most nonmetallic elements achieve noble-gas electron configurations by gaining electrons to become _____, or negatively charged ions.

A. anions
B. cations
C. isotopes
D. neutral atoms

Question 7 [Practice]

How many valence electrons are available in NO_2^-?

A. 16
B. 19
C. 18
D. 17

Question 8 [Practice]

How many valence electrons does NO_2 have?

A. 18
B. 16
C. 15
D. 17

Question 9 [Practice]

Which molecule has an odd number of valence electrons?

A. ClO_2
B. CO_2
C. H_2O
D. CH_4

Question 10 [Practice]

In NO_2, how many electrons are around the central N atom?

A. 10
B. 7
C. 8
D. 6

Question 11 [Practice]

How many electrons does boron share in BH_3?

A. 6
B. 8
C. 3
D. 4

Question 12 [Practice]

What is the configuration of Zn^{2+}?

A. $[Ar]4s^2$
B. $[Ar]4s^23d^8$
C. $[Ar]3d^{10}$
D. $[Ar]4s^13d^9$

Question 13 [Practice]

When zinc forms an ion, which electrons are lost first?

A. 3d
B. 3p
C. 4s
D. 4p

Question 14 [Practice]

A pseudo-noble-gas configuration has an outer energy level with full:

A. d only
B. s, p, d and f
C. s, p and d sublevels
D. s and p only

Question 15 [Practice]

Cadmium is $[Kr]5s^2 4d^{10}$. What is its pseudo-noble-gas ion configuration?

A. $[Ar]4d^{10}$
B. $[Kr]4d^{10}$
C. $[Kr]5s^2 4d^8$
D. $[Kr]5s^2$

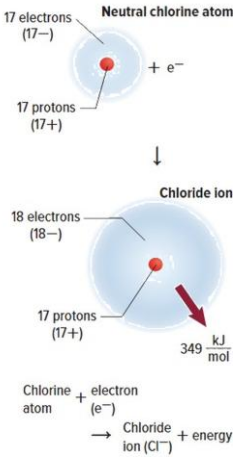
Question 16 [Practice]

Which ion has a pseudo-noble-gas configuration?

A. Sc^{3+}
B. Cl^-
C. Cu^+
D. Na^+

Q14 Learning outcome: Ion formation

Textbook: p. 66 – Figure 4

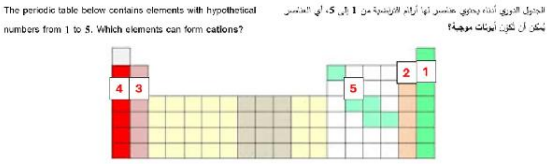


Key points from the lesson

- Nonmetals (right side of the periodic table) easily gain electrons to attain a stable noble-gas configuration.
- Chlorine $1s^2 2s^2 2p^6 3s^2 3p^5$ gains one electron → chloride ion Cl^- $1s^2 2s^2 2p^6 3s^2 3p^6$ (same as argon).
- Forming Cl^- releases energy: 349 kJ/mol (Figure 4). Cl^- has 17 protons and 18 electrons.
- An anion is a negatively charged ion. Name: root + -ide (chlorine → chloride, sulfur → sulfide, oxygen → oxide).
- Group 15 gains 3 e^- (3−), group 16 gains 2 e^- (2−), group 17 gains 1 e^- (1−).

Question 1 [Past exam 2025–26]

The periodic table contains elements with hypothetical numbers 1 to 5. Which elements can form cations?



A. 3 and 4
B. 1 and 3
C. 2 and 4
D. 2 and 5

Question 2 [Mock exam]

The ion formed by group 17 elements has a charge of:

A. 1+
B. 2−
C. 1−
D. 7+

Question 3 [Mock exam]

A chlorine atom gains one electron to become Cl^-. Its electron arrangement becomes like the noble gas:

A. Ne
B. He
C. Kr
D. Ar

Question 4 [Practice]

Which statement about positive ions is INCORRECT?

A. they are also known as cations
B. they are formed when electrons are removed from atoms
C. they are larger than the atom from which they were formed
D. they are smaller than the atom from which they were formed

Question 5 [Practice]

What is the formula of the compound formed between strontium ions and nitrogen ions?

A. SrN
B. Sr_3N_2
C. Sr_2N_3
D. SrN_2

Question 6 [Practice]

What is the electron configuration of the copper(I) ion, Cu^+?

A. $[Ar]3d^{10}$
B. $[Ar]3d^9$
C. $[Ar]3d^{10}4s^1$
D. $[Ar]3d^94s^1$

Question 7 [Practice]

Which statement about negative ions is INCORRECT?

A. they are also known as anions
B. they are formed when atoms gain electrons
C. they are larger than the atom from which they were formed
D. they are smaller than the atom from which they were formed

Question 8 [Practice]

Which statement best describes the force of attraction in a chemical bond?

A. It is an attraction between oppositely charged ions, or between nuclei and shared valence electrons.
B. It is an attraction between two nuclei only.
C. It is a repulsion between valence electrons of different atoms.
D. It is an attraction between neutrons and electrons.

Question 9 [Practice]

Which pair of atoms is most likely to form an ionic bond?

A. Na and F
B. C and F
C. N and F
D. O and F

Question 10 [Practice]

How does the reactivity of group 18 elements compare with that of group 17 elements?

A. Group 18 elements are almost unreactive because they have a full octet, while group 17 elements are very reactive because they need only one electron.
B. Group 18 elements are more reactive because they have more valence electrons.
C. Both groups are equally reactive because they are nonmetals.
D. Group 17 elements are unreactive because they have seven valence electrons.

Question 11 [Practice]

The structure normally associated with ionic bonding is…

A. a giant lattice
B. a simple molecule
C. a giant molecule
D. positive ions surrounded by a sea of electrons

Question 12 [Practice]

Why are alkali metals the most reactive metals?

A. They have only one valence electron, which is easily lost to reach a noble-gas configuration.
B. They have seven valence electrons and gain one easily.
C. They have a full octet of valence electrons.
D. They have the highest ionization energies.

Question 13 [Practice]

Sulfur ([Ne]$3s^2 3p^4$) becomes stable by:

A. losing six electrons
B. gaining two electrons → S^{2-}
C. gaining six electrons
D. losing two electrons → S^{2+}

Question 14 [Practice]

The chloride ion has the same configuration as:

A. Kr
B. Ar
C. Na^+
D. Ne

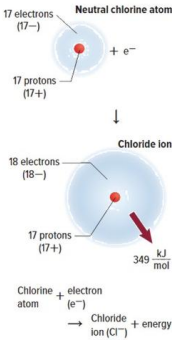
Question 15 [Practice]

What is the name of the ion N^{3-}?

A. nitrate
B. nitride
C. nitrogen
D. nitrite

Question 16 [Practice]

According to Figure 4, forming one mole of Cl^- from Cl atoms:



A. absorbs 349 kJ
B. releases 349 kJ
C. involves no energy change
D. absorbs 498 kJ

Question 17 [Practice]

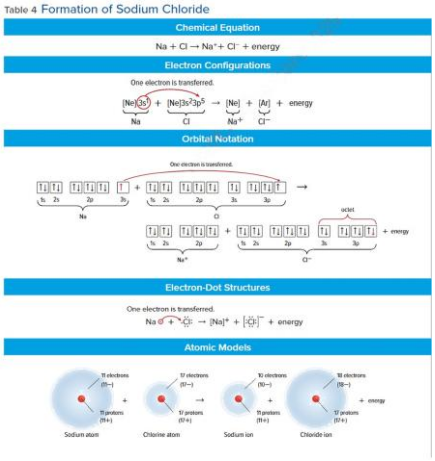
Which group's elements gain two electrons to form 2− ions?

A. Group 15
B. Group 16
C. Group 2
D. Group 17

Q15

Learning outcome: Describe how ions (cations and anions) form to fulfill the octet rule

Textbook: p. 69 – Table 4

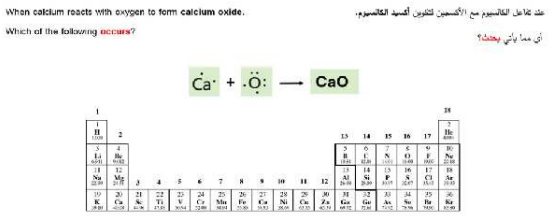


Key points from the lesson

- NaCl is a binary ionic compound: it contains only two elements – a metallic cation and a nonmetallic anion (MgO is another example).
- One electron is transferred from Na to Cl: Na + Cl → Na^+ + Cl^- + energy.
- Electron configurations: [Ne]$3s^1$ + [Ne]$3s^2 3p^5$ → [Ne] (Na^+) + [Ar] (Cl^-).
- Electron-dot: Na· + ·Cl (7 dots) → $[Na]^+$ [Cl with 8 dots]$^-$ – both ions now have stable octets.
- Atomic models: Na (11e^-) → Na^+ (10e^-) smaller; Cl (17e^-) → Cl^- (18e^-) larger. Overall the process releases energy.

Question 1 [Past exam 2025–26]

When calcium reacts with oxygen to form calcium oxide (Ca· + ·Ö: → CaO), which of the following occurs?



A. The calcium atom loses 2 electrons and attains the electron configuration of argon (Ar)
B. The oxygen atom gains 1 electron and attains the electron configuration of neon (Ne)
C. The calcium atom loses 2 electrons and becomes Ca^{2-}
D. The oxygen atom gains 2 electrons and becomes O^{2+}

Question 2 [Mock exam]

Using electron-dot structures, what is the formula of the ionic compound formed from calcium, Ca (group 2), and chlorine, Cl (group 17)?

A. $CaCl_2$
B. CaCl
C. Ca_2Cl
D. Ca_2Cl_3

Question 3 [Practice]

Zinc forms Zn^{2+}. Which configuration does Zn^{2+} have?

A. [Ar]$4s^2 3d^8$
B. [Ar]$3d^{10}$
C. [Ar]$4s^2$
D. [Ne]$3s^2 3p^6$

Question 4 [Practice]

How many electrons are in the outer levels of an ion with a pseudo-noble gas configuration?

A. 8
B. 18
C. 2
D. 10

Question 5 [Practice]

Metals form which type of ions?

A. negative
B. anions
C. cations
D. neutral

Question 6 [Practice]

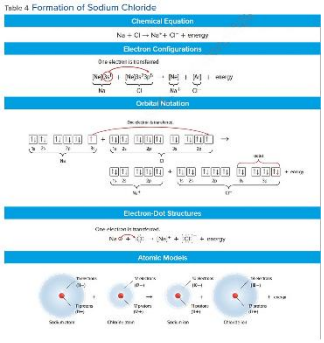
Rule 1: in the name of an ionic compound…

A. the anion is named first
B. the cation is named first, followed by the anion
C. the larger ion is named first
D. the order does not matter

Question 7 [Practice]

How many electrons are transferred when one formula unit of NaCl forms?

A. 1
B. 2
C. 7
D. 8



Question 8 [Practice]

In Table 4, the electron configuration of Na^+ is written as:

A. $[Ne]3s^2$
B. [Ne]
C. [Ar]
D. $[Ne]3s^1$

Question 9 [Practice]

In the electron-dot structure of the chloride ion, how many dots surround Cl?

A. 1
B. 6
C. 7
D. 8

Question 10 [Practice]

Which is a binary ionic compound?

A. CO_2
B. MgO
C. $NaNO_3$
D. H_2O

Question 11 [Practice]

Compared with the atoms, in NaCl formation:

A. sizes do not change
B. Na^+ is larger than Na and Cl^- is smaller than Cl
C. Na^+ is smaller than Na and Cl^- is larger than Cl
D. both ions are smaller

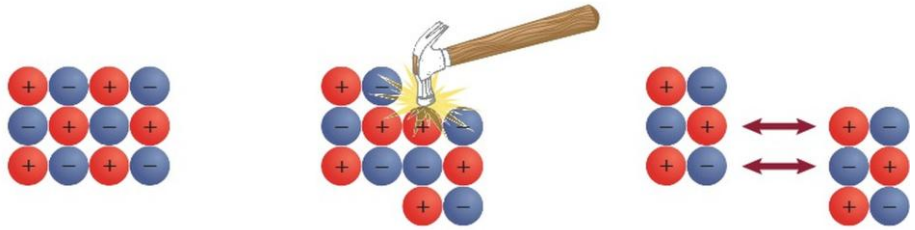
Q16

Learning outcome: Explain the physical properties of ionic compounds: melting and boiling point, conductivity (solid, molten, aqueous) and solubility

Textbook: pp. 72–73 – Table 5 + Figure 8

Table 5 Melting and Boiling Points of Some Ionic Compounds

Compound	Melting Point (°C)	Boiling Point (°C)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
$CaCl_2$	782	>1600
NaCl	801	1413
MgO	2852	3600



Undisturbed ionic crystal
Before the force is applied, the crystal has a uniform pattern of ions.

Applied force realigns particles.
If the applied force is strong enough, it pushes the ions out of alignment.

Forces of repulsion break crystal apart.
A repulsive force created by nearby like charged ions breaks apart the crystal.

Key points from the lesson

- Melting point, boiling point and hardness depend on how strongly the particles are held together.
- Ionic bonds are strong, so ionic crystals need a large amount of energy to break apart → high melting and boiling points (NaCl: 801 °C; MgO: 2852 °C).
- Ions with larger charges (Mg^{2+}, O^{2-}) attract more strongly → higher melting point.
- Ionic crystals are hard, rigid and brittle: a strong force shifts the ions so like charges face each other; repulsion breaks the crystal (Figure 8).
- Many ionic crystals (gemstones) are brightly colored because of transition metals in the lattice.

Question 1 [Past exam 2025–26]

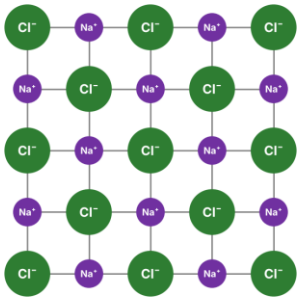
What explains the breaking apart of the ionic crystal when struck (figure)?



A. There are strong attractive forces that hold the ions in place, and the applied force is enough to overcome the attraction
B. There are strong attractive forces that hold the ions in place, but the applied force is not enough to overcome the attraction
C. There are weak attractive forces between the ions, and the applied force is enough to overcome the attraction
D. There are weak attractive forces between the ions, and the applied force is not enough to overcome the attraction

Question 2 [Mock exam]

The figure shows how the ions are arranged in sodium chloride, NaCl. This arrangement is called a:



A. Molecule
B. Crystal lattice
C. Covalent bond
D. Electron cloud

Question 3 [Mock exam]

Which is a physical property of ionic compounds?

A. Very low melting point
B. Conduct electricity as solids
C. High melting point and conduct electricity when dissolved in water
D. Never dissolve in water

Question 4 [Practice]

Which compound has the highest melting point?

A. NaCl
B. KBr
C. MgO
D. NaI

Question 5 [Practice]

Which property of ionic compounds is caused by applying an external force?

A. brittleness
B. high melting point
C. solubility in water
D. conductivity when molten

Question 6 [Practice]

I can hit a metal with a hammer without it shattering because of its…

A. ductility
B. malleability
C. conductivity
D. luster

Question 7 [Practice]

Hardness, boiling point and melting point are _____ properties of ionic compounds.

A. physical
B. chemical
C. nuclear
D. electrical

Question 8 [Practice]

Using Table 5, which compound has the highest melting point?

Table 5 Melting and Boiling Points of Some Ionic Compounds

Compound	Melting Point (°C)	Boiling Point (°C)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
$CaCl_2$	782	>1600
NaCl	801	1413
MgO	2852	3600

A. NaCl
B. KBr
C. NaI
D. MgO

Question 9 [Practice]

Ionic compounds have high melting points because:

A. the attraction between oppositely charged ions is strong
B. their molecules are large
C. they are gases
D. they contain delocalized electrons

Question 10 [Practice]

Why does an ionic crystal shatter when struck hard (Figure 8)?

Table 5 Melting and Boiling Points of Some Ionic Compounds

Compound	Melting Point (°C)	Boiling Point (°C)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
$CaCl_2$	782	>1600
NaCl	801	1413
MgO	2852	3600

A. Covalent bonds break.
B. The ions melt.
C. Like-charged ions are pushed next to each other and repel.
D. Electrons are removed.

Question 11 [Practice]

Ionic crystals are best described as:

A. good conductors as solids
B. ductile
C. hard, rigid and brittle
D. soft and malleable

Question 12 [Practice]

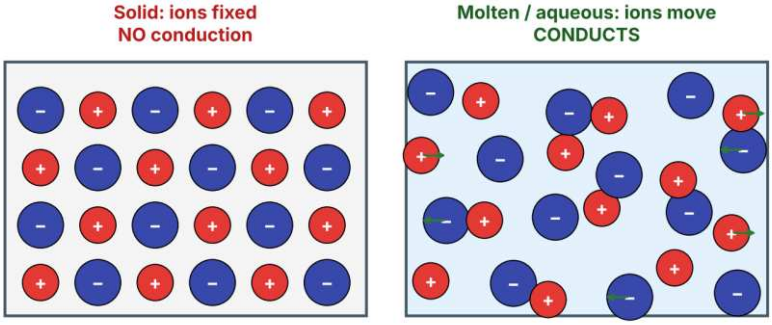
Why is the melting point of MgO much higher than that of NaCl?

A. Mg^{2+} and O^{2-} have larger charges, so the attraction is stronger.
B. MgO is a molecule.
C. MgO has more atoms.
D. NaCl is a metal.

Q17

Learning outcome: Explain the physical properties of ionic compounds: melting and boiling point, conductivity and solubility

Textbook: p. 73



Key points from the lesson

- Conduction requires freely moving charged particles. In the solid state, ions are locked in fixed positions → ionic solids do NOT conduct electricity.
- When melted (liquid) or dissolved in water, the ions are free to move → excellent conductors.
- Electrolyte: an ionic compound whose aqueous solution conducts an electric current.
- Forming an ionic compound is always exothermic. Lattice energy: the energy required to separate 1 mol of the ions of an ionic compound.
- Greater lattice energy → stronger attraction. Smaller ions → greater lattice energy (a Li compound > a K compound with the same anion).

Question 1 [Past exam 2025–26]

Which is correct about the properties of ionic compounds? (1) Ionic crystals are brittle, hard, rigid solids. (2) Their melting and boiling points are high. (3) They conduct electric current when melted or dissolved in water.

A. 1, 2 and 3
B. 1 only
C. 1 and 3
D. 2 only

Question 2 [Mock exam]

When an ionic crystal is struck with a strong force, it:

A. Bends without breaking
B. Stretches into a wire
C. Becomes an electrical conductor
D. Shatters because like-charged ions line up and repel

Question 3 [Practice]

Predict the physical state of astatine (At_2, below iodine) at room temperature.

A. Solid
B. Gas
C. Liquid
D. Plasma

Question 4 [Practice]

In a metallic lattice, how many other metal atoms closely surround each metal atom?

A. 2 to 4
B. 4 to 6
C. 8 to 12
D. 20 to 30

Question 5 [Practice]

Why are the melting points of metals less extreme than their boiling points?

A. Cations and electrons are mobile, so moving past each other needs little energy; boiling needs atoms completely separated
B. Metals have no bonds in the liquid
C. Boiling needs less energy than melting
D. All electrons leave the metal when it melts

Question 6 [Practice]

Which metal is a liquid at room temperature and is used in thermometers?

A. Tungsten
B. Mercury
C. Copper
D. Silver

Question 7 [Practice]

Solid NaCl does not conduct electricity because:

A. it has no ions
B. its ions are locked in fixed positions
C. its electrons move freely
D. it is a molecule

Question 8 [Practice]

In which states does KBr conduct electricity?

A. solid and molten
B. never
C. solid only
D. molten and aqueous

Question 9 [Practice]

An electrolyte is:

A. any solid crystal
B. an ionic compound whose aqueous solution conducts electricity
C. a metal that conducts heat
D. a covalent gas

Question 10 [Practice]

Which compound has the greater lattice energy?

A. Neither has lattice energy
B. LiF, because Li^+ is smaller
C. KF, because K^+ is larger
D. They are equal

Question 11 [Practice]

Lattice energy is the energy:

A. released when an electron is added
B. required to separate 1 mol of the ions of an ionic compound
C. required to melt a metal
D. of a covalent bond

Q18

Learning outcome: Write the chemical formula (stock method) of an ionic compound containing monoatomic and polyatomic ions

Textbook: pp. 75–78 – Tables 7, 8 + Examples 1, 2

Table 7 Common Monatomic Ions

Group	Atoms that Commonly Form Ions	Charge of Ions
1	H, Li, Na, K, Rb, Cs	1+
2	Be, Mg, Ca, Sr, Ba	2+
15	N, P, As	3−
16	O, S, Se, Te	2−
17	F, Cl, Br, I	1−

Table 8 Monatomic Metal Ions

Group	Common Ions
3	Sc^{3+}, Y^{3+}, La^{3+}
4	Ti^{2+}, Ti^{3+}
5	V^{2+}, V^{3+}
6	Cr^{2+}, Cr^{3+}
7	Mn^{2+}, Mn^{3+}, Tc^{2+}
8	Fe^{2+}, Fe^{3+}
9	Co^{2+}, Co^{3+}
10	Ni^{2+}, Pd^{2+}, Pt^{2+}, Pt^{4+}
11	Cu^{+}, Cu^{2+}, Ag^{+}, Au^{+}, Au^{3+}
12	Zn^{2+}, Cd^{2+}, Hg_2^{2+}, Hg^{2+}
13	Al^{3+}, Ga^{2+}, Ga^{3+}, In^{+}, In^{2+}, In^{3+}, Tl^{+}, Tl^{3+}
14	Sn^{2+}, Sn^{4+}, Pb^{2+}, Pb^{4+}

Key points from the lesson

- Formula unit: the simplest ratio of ions in an ionic compound. Its total charge is zero: $MgCl_2$ = one Mg^{2+} + two Cl^{-}.
- Table 7: group 1 → 1+, 2 → 2+, 15 → 3−, 16 → 2−, 17 → 1−. Transition metals and groups 13–14 can form more than one ion (Table 8).
- Oxidation number of a monatomic ion = its charge (Na^{+}: 1+, Cl^{-}: 1−).
- Write the cation first, then the anion; use subscripts so that electrons lost = electrons gained.
- Example 1: $K^{+} + O^{2-} \rightarrow K_2O$. Example 2: $Al^{3+} + S^{2-} \rightarrow Al_2S_3$ (6 electrons transferred).

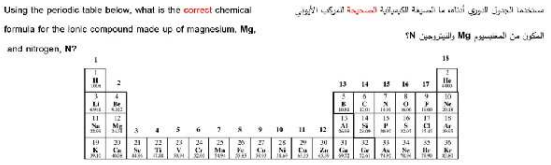
Question 1 [Past exam 2025–26]

What is the correct chemical formula for the ionic compound formed by Ba^{2+} and PO_4^{3-}?

A. $Ba_3(PO_4)_2$
B. $Ba_2(PO_4)_3$
C. Ba_3PO_4
D. $BaPO_4$

Question 2 [Past exam 2025–26]

Using the periodic table, what is the correct chemical formula for the ionic compound made of magnesium (Mg) and nitrogen (N)?



A. Mg_3N_2
B. MgN_2
C. MgN
D. Mg_2N_3

Question 3 [Practice]

The formula of the compound formed from ammonium (NH_4^+) and oxide (O^{2-}) is…

A. NH_4O
B. N_2H_8O
C. $(NH_4)_2O$
D. NH_4O_2

Question 4 [Practice]

Book Q38: name and formula of the compound of sodium ions and nitrite ions

A. sodium nitrate, $NaNO_3$
B. sodium nitrite, $NaNO_2$
C. sodium nitride, Na_3N
D. sodium nitrite, $Na(NO_2)_2$

Question 5 [Practice]

Booklet Q123: in the chemical formula for an ionic compound, which item is written first?

A. positive ion
B. negative ion
C. subscript
D. charge

Question 6 [Practice]

What is the formula of the compound formed from Ca^{2+} and Cl^-?

A. $CaCl$
B. Ca_2Cl
C. $CaCl_2$
D. Ca_2Cl_3

Question 7 [Practice]

What is the formula of aluminum oxide?

A. Al_2O_3
B. AlO_3
C. Al_3O_2
D. AlO

Question 8 [Practice]

Magnesium and nitrogen form a compound with the formula:

A. Mg_2N_3
B. Mg_3N_2
C. MgN
D. MgN_2

Question 9 [Practice]

What is the oxidation number of oxygen in K_2O?

A. 1+
B. 2+
C. 2−
D. 1−

Question 10 [Practice]

Cesium and nitrogen form a compound with the formula:

A. Cs_3N
B. CsN_3
C. CsN
D. Cs_2N_3

Q19

Learning outcome: Write the chemical formula (stock method) of an ionic compound containing monoatomic and polyatomic ions

Textbook: pp. 79–80 – Tables 10, 11

Table 9 Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$C_2H_3O_2^-$	acetate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	CO_3^{2-}	carbonate
CN^-	cyanide	SO_3^{2-}	sulfite
MnO_4^-	permanganate	SO_4^{2-}	sulfate
HCO_3^-	hydrogen carbonate	$S_2O_3^{2-}$	thiosulfate
ClO^-	hypochlorite	O_2^{2-}	peroxide
ClO_2^-	chlorite	CrO_4^{2-}	chromate
ClO_3^-	chlorate	$Cr_2O_7^{2-}$	dichromate
ClO_4^-	perchlorate	HPO_4^{2-}	hydrogen phosphate
BrO_3^-	bromate	PO_4^{3-}	phosphate
IO_3^-	iodate	AsO_4^{3-}	arsenate

Key points from the lesson

- A polyatomic ion acts as one unit with one charge. Never change its subscripts; use parentheses for more than one: $(NH_4)_2O$, $Ca_3(PO_4)_2$.
- Oxyanions: more oxygen → -ate, fewer → -ite (NO_3^- nitrate, NO_2^- nitrite; SO_4^{2-} sulfate, SO_3^{2-} sulfite).
- Chlorine: ClO_4^- perchlorate, ClO_3^- chlorate, ClO_2^- chlorite, ClO^- hypochlorite.
- Naming: cation first, then anion; monatomic anion = root + -ide (CsBr = cesium bromide; NaOH = sodium hydroxide).
- Stock method: for metals with more than one charge, write the oxidation number as a Roman numeral: FeO iron(II) oxide, Fe_2O_3 iron(III) oxide.

Question 1 [Past exam 2025–26]

What is the correct name for the ionic compound $NaIO_4$?

A. Sodium periodate
B. Sodium iodate
C. Sodium iodite
D. Sodium hypoiodite

Question 2 [Practice]

When more than one polyatomic ion is needed in a formula, you should… ةعئاشلا تارذلا ةديدع تانوياً Common polyatomic ions · NH_4^+ CH_3COO^- Ammonium Acetate NO_2^- SO_3^{2-} Nitrite Sulfite NO_3^- SO_4^{2-} Nitrate Sulfate OH^- CO_3^{2-} Hydroxide Carbonate CN^- CrO_4^{2-} Cyanide Chromate MnO_4^- $Cr_2O_7^{2-}$ Permanganate Dichromate HCO_3^- HPO_4^{2-} Hydrogen carbonate Hydrogen phosphate ClO^- $C_2O_4^{2-}$ Hypochlorite Oxalate ClO_2^- O_2^{2-} Chlorite Peroxide ClO_3^- PO_4^{3-} Chlorate Phosphate ClO_4^- PO_3^{3-} Perchlorate Phosphite $H_2PO_4^-$ Dihydrogen phosphate

A. change the subscripts inside the ion
B. place the ion in parentheses and write the subscript outside
C. write the ion twice
D. add a Roman numeral

Question 3 [Practice]

Book Q36: ion X is 2+ and ion Y is 1−. The formula unit is…

A. XY
B. X_2Y
C. XY_2
D. X_2Y_2

Question 4 [Practice]

Book Q37: name and formula of the compound formed from Mg and Cl

A. magnesium chloride, $MgCl_2$
B. magnesium chlorine, MgCl
C. magnesium dichloride, Mg_2Cl
D. manganese chloride, $MnCl_2$

Question 5 [Practice]

An oxyanion is a polyatomic ion composed of...

Use the figure above

A. a metal bonded to hydrogen
B. an element, usually a nonmetal, bonded to one or more oxygen atoms
C. two metals
D. oxygen only

Question 6 [Practice]

What is the formula of calcium phosphate?

A. $Ca_3(PO_4)_2$
B. Ca_3PO_4
C. $Ca_2(PO_4)_3$
D. $CaPO_4$

Question 7 [Practice]

What is the name of ClO_3^-?

A. hypochlorite
B. chlorite
C. chlorate
D. perchlorate

Question 8 [Practice]

What is the Stock name of Fe_2O_3?

A. iron oxide
B. iron(III) oxide
C. diiron trioxide
D. iron(II) oxide

Question 9 [Practice]

What is the formula of aluminum carbonate?

A. $AlCO_3$
B. $Al_3(CO_3)_2$
C. $Al_2(CO_3)_3$
D. $Al(CO_3)_3$

Question 10 [Practice]

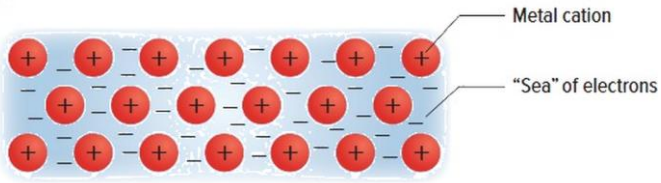
What is the name of $Cu(NO_3)_2$?

A. copper(II) nitrate
B. copper(I) nitrate
C. copper nitride
D. copper nitrite

Q20

Learning outcome: Describe the formation of metallic bonds

Textbook: p. 82 – Figure 10



Key points from the lesson

- In the solid state, metals form lattices in which 8 to 12 other metal atoms closely surround each atom.
- Metal atoms do not share or lose their valence electrons to neighbors; instead their outer energy levels overlap.
- Electron sea model: all metal atoms contribute their valence electrons to a "sea" of electrons surrounding the metal cations.
- These electrons are not held by any specific atom and move freely: delocalized electrons.
- Metallic bond: the attraction of a metallic cation for delocalized electrons.

Question 1 [Past exam 2025–26]

Regarding the figure (metal cations in a sea of electrons), which is correct? (1) The outer electrons of the bonded metal atoms are held by a specific atom. (2) They are not held by a specific atom. (3) They can move easily from one atom to the next.

Regarding the figure below, which of the following is correct? فيما يتعلق بالشكل أدناه، أي مما يلي صحيح؟



1	ترتبط الإلكترونات الموجودة في مستويات الطاقة الخارجية للذرات الفلزية المترابطة بذرة معينة The electrons present in the outer energy levels of the bonding metallic atoms are held by a specific atom
2	لا ترتبط الإلكترونات الموجودة في مستويات الطاقة الخارجية للذرات الفلزية المترابطة بذرة معينة The electrons present in the outer energy levels of the bonding metallic atoms are not held by a specific atom
3	يمكن للإلكترونات الموجودة في مستويات الطاقة الخارجية للذرات الفلزية الانتقال بسهولة من ذرة إلى أخرى The electrons present in the outer energy levels of the bonding metallic atoms can move easily from one atom to the next

A. 2 and 3
B. 1 and 3
C. 1 only
D. 2 only

Question 2 [Practice]

Metals are malleable because, when struck, one plane of metal atoms…

A. can slide past another plane without breaking bonds
B. cannot easily move out of the way
C. moves in a way that maximizes repulsive forces
D. bonds to the plane directly beneath it

Question 3 [Practice]

The arrangement of valence electrons in a metallic bond is best described as…

A. fixed positions in a lattice
B. a sea of free-moving electrons
C. concentrated electron density around specific atoms
D. electron pairs existing in multiple bonds

Question 4 [Practice]

In metallic bonding…

A. one atom takes the outer-shell electrons from another atom
B. positively charged areas of one atom bond to negatively charged areas of another
C. a couple of atoms share their electrons with each other
D. some electrons are shared by all the atoms in the material

Question 5 [Practice]

Under typical conditions, how many electrons is the maximum gained in the formation of anions?

A. 1
B. 2
C. 3
D. 4

Question 6 [Practice]

An ionic bond is…

A. the electrostatic attraction between oppositely charged ions
B. a shared pair of electrons between two atoms
C. the attraction between two neutral atoms
D. a bond between two metal atoms

Question 7 [Practice]

The attractive forces that hold particles together in ionic, covalent and metallic bonds are called…

A. intramolecular forces
B. intermolecular forces
C. dispersion forces
D. hydrogen bonds

Question 8 [Practice]

Single covalent bonds are also called…

A. pi bonds
B. sigma bonds
C. ionic bonds
D. metallic bonds

Question 9 [Practice]

Why is sugar ($C_{12}H_{22}O_{11}$), a molecular compound, a solid at room temperature?

A. Its large molar mass lets many weak attractions combine
B. It contains ionic bonds
C. It has a sea of electrons
D. It is a covalent network

Question 10 [Practice]

According to the electron sea model, the valence electrons in a metal are:

A. delocalized, moving freely around the metal cations
B. transferred to nonmetal atoms
C. shared in pairs between two neighbouring atoms
D. held tightly in the inner energy levels

Question 11 [Practice]

A metallic bond is the attraction between:

A. two nuclei and a shared pair
B. metallic cations and delocalized electrons
C. neighboring molecules
D. a cation and an anion

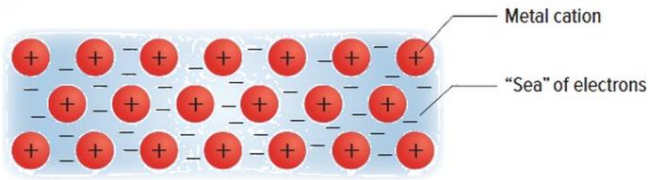
Question 12 [Practice]

Electrons in a metal are called delocalized because they:

A. move freely and are not held by a specific atom
B. are held by one atom
C. are transferred to a nonmetal
D. are in the nucleus

Question 13 [Practice]

In Figure 10, the circles with + signs represent:



A. neutral atoms
B. protons
C. anions
D. metal cations

Question 14 [Practice]

In a metallic lattice, each atom is closely surrounded by:

A. 1–2 atoms
B. 3–4 atoms
C. 8–12 atoms
D. 20–30 atoms

Question 15 [Practice]

Which model describes bonding in metals?

A. electron sea model
B. Bohr model
C. Lewis structure
D. resonance model

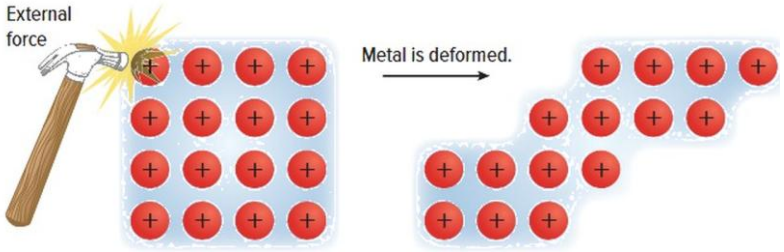
Q21

Learning outcome: Explain some physical properties of metals

Textbook: pp. 83–84 – Figure 11 + Table 12

Table 12 Melting and Boiling Points

Element	Melting Point (°C)	Boiling Point (°C)
Lithium	180	1342
Tin	232	2602
Aluminum	660	2519
Barium	727	1897
Silver	962	2162
Copper	1085	2562



Key points from the lesson

• Melting points vary greatly (Hg is liquid; W melts at 3422 °C). Boiling points are much higher because atoms must be completely separated from the lattice.

• Thermal and electrical conductivity: mobile delocalized electrons carry heat and current; they also interact with light → luster.

• Malleable (hammered into sheets) and ductile (drawn into wire): cations are pushed past each other through the electron sea (Figure 11). Metals are durable.

• Hardness and strength increase with the number of delocalized electrons: transition metals (Cr, Fe, Ni) are strong (d electrons are delocalized); alkali metals (ns^1) are soft.

Question 1 [Past exam 2025–26]

Which of the following explains the heat and electrical conductivity of metals?

A. The movement of delocalized electrons around positive metallic cations
B. The strong attraction of metallic cations to the delocalized electrons
C. The delocalized electrons interact with light, absorbing and releasing photons
D. The movement of delocalized electrons around negative anions

Question 2 [Practice]

Which property is NOT explained by metallic bonding?

A. electrical conductivity
B. thermal conductivity
C. brittleness
D. ductility

Question 3 [Practice]

What is responsible for the good electrical conductivity of metals?

A. the arrangement of metal atoms in separate layers
B. the high density of metal atoms in the lattice
C. the ability of electrons to move freely about the crystal structure
D. many orbitals separated by small energies

Question 4 [Practice]

In general, what can be said of the melting points of metals?

A. They are low
B. They are high
C. They are lower than nonmetals
D. They do not have melting points

Question 5 [Practice]

Why do solid ionic compounds not conduct electricity, while molten or dissolved ionic compounds do?

A. In the solid, the ions are locked in fixed positions; when melted or dissolved, the ions are free to move.
B. In the solid, the electrons are delocalized; when melted, they become fixed.
C. Solids contain no ions; ions form only when melted.
D. Water molecules carry the electric current, not the ions.

Question 6 [Practice]

What do metals conduct?

A. heat only
B. electricity only
C. both heat and electricity
D. neither

Question 7 [Practice]

Which one of the following is NOT true of metallic bonding?

A. It gives rise to excellent electrical conductivity.
B. Electrons are free to move throughout the structure.
C. The strength of metallic bonds increases down a group.
D. The strength of metallic bonding affects the boiling point of metals.

Question 8 [Practice]

Use Table 12. Which metal has the highest melting point?

A. Lithium (180 °C)
B. Tin (232 °C)
C. Aluminum (660 °C)
D. Copper (1085 °C)

Question 9 [Practice]

Consider the statements: (1) Ionic compounds are electrolytes that conduct electricity in the liquid (molten) state. (2) Ionic compounds are crystalline solids. (3) Ionic compounds are not soluble in water. Which statements are true?

A. 3 only
B. Only 2 and 3
C. Only 1 and 3
D. Only 1 and 2

Question 10 [Practice]

The luster (shine) of metals is caused by…

A. delocalized electrons absorbing and releasing photons
B. cations reflecting heat
C. metal atoms losing protons
D. the metal being an ionic compound

Question 11 [Practice]

Metals are good conductors of electricity because:

A. they share electron pairs
B. their ions are fixed
C. they have high melting points
D. their delocalized electrons move freely

Question 12 [Practice]

A metal that can be drawn into a wire is described as:

A. malleable
B. brittle
C. lustrous
D. ductile

Question 13 [Practice]

Why are boiling points of metals much higher than their melting points (Table 12)?

Table 12 Melting and Boiling Points

Element	Melting Point (°C)	Boiling Point (°C)
Lithium	180	1342
Tin	232	2602
Aluminum	660	2519
Barium	727	1897
Silver	962	2162
Copper	1085	2562

A. Boiling requires atoms to be completely separated from the lattice.
B. Metals lose electrons when boiling.
C. Melting breaks covalent bonds.
D. Cations become anions.

Question 14 [Practice]

Iron is much harder than sodium because iron:

A. has more delocalized electrons (including d electrons)
B. forms ionic bonds
C. has fewer electrons
D. is a nonmetal

Question 15 [Practice]

When a metal is struck by a hammer (Figure 11), it:

Table 12 Melting and Boiling Points

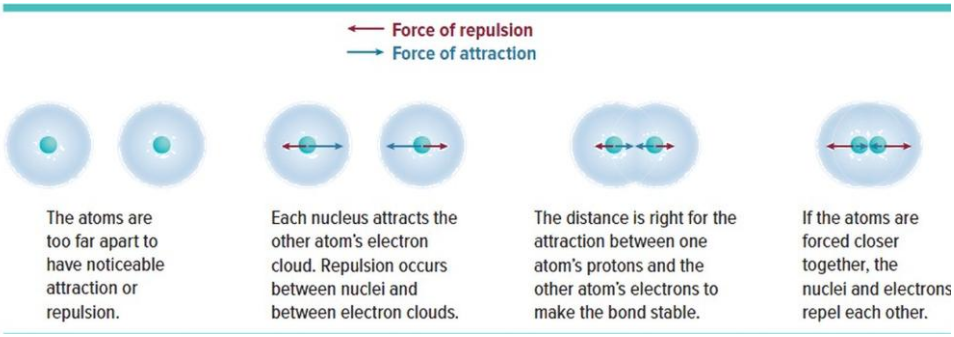
Element	Melting Point (°C)	Boiling Point (°C)
Lithium	180	1342
Tin	232	2602
Aluminum	660	2519
Barium	727	1897
Silver	962	2162
Copper	1085	2562

A. loses all its electrons
B. shatters like an ionic crystal
C. deforms because cations move through the electron sea
D. melts immediately

Q22

Learning outcome: Define a covalent bond using an energy diagram

Textbook: p. 92 – Figure 2



Key points from the lesson

- A covalent bond is the chemical bond that results from sharing valence electrons. A molecule forms when two or more atoms bond covalently.
- Covalent bonds generally form between nonmetals. Diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.
- As two F atoms approach: attraction (nucleus – other atom's electrons) and repulsion (electron – electron, nucleus – nucleus).
- The bond forms at the distance of maximum net attraction (lowest potential energy). Each F now has an octet by sharing one pair.
- If the nuclei move closer, repulsion increases and exceeds attraction.

Question 1 [Practice]

How does water vapor condense on a cold surface?

A. Vapor molecules lose energy to the cold surface, slow down and form hydrogen bonds → liquid
B. The cold surface gives energy to the vapor
C. Water molecules split into H and O
D. The vapor turns into a solid without losing energy

Question 2 [Practice]

Electronegativity is…

A. the energy needed to remove an electron
B. the energy released when an atom gains an electron
C. the size of an atom
D. the relative ability of an atom to attract the electrons in a chemical bond

Question 3 [Practice]

Bond-dissociation energy is always…

A. a positive value
B. a negative value
C. zero
D. equal to the bond length

Question 4 [Practice]

A reaction in which the energy released by forming bonds in the products is greater than the energy needed to break the bonds in the reactants is _____.

A. exothermic
B. endothermic
C. neutral
D. reversible

Question 5 [Practice]

Burning charcoal (Figure 11) is exothermic because...

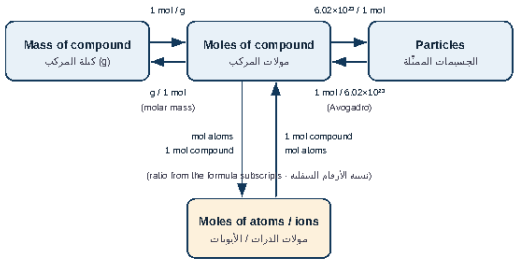


Figure 11 · Mole relationships from a chemical formula · علاقات المول من الصيغة

A. more energy is released forming the bonds in CO_2 than is needed to break C–C and O–O bonds
B. no bonds are broken
C. breaking bonds releases energy
D. charcoal has no bonds

Question 6 [Practice]

Using the bond energy C=O = 799 kJ/mol, what is the total energy needed to break all the bonds in CO_2 (O=C=O)?

A. 1598 kJ/mol
B. 799 kJ/mol
C. 2397 kJ/mol
D. 1197 kJ/mol

Question 7 [Practice]

Why do noble gases seldom form compounds?

A. They have high potential energy
B. They have stable electron arrangements with a full outer energy level
C. They have no electrons
D. They are metals

Question 8 [Practice]

Using the bond energy O–H = 463 kJ/mol, what is the total energy needed to break all the bonds in a water molecule, H–O–H?

A. 926 kJ/mol
B. 463 kJ/mol
C. 1389 kJ/mol
D. 232 kJ/mol

Question 9 [Practice]

Which statement about stability is correct?

A. Higher energy states are more stable
B. Energy is not related to stability
C. Separated atoms always have the lowest energy
D. Lower energy states are more stable

Question 10 [Practice]

What is the relationship between bond length and bond-dissociation energy? (Booklet Q31)

A. Inverse: the shorter the bond, the greater the energy
B. Direct: the longer the bond, the greater the energy
C. No relationship

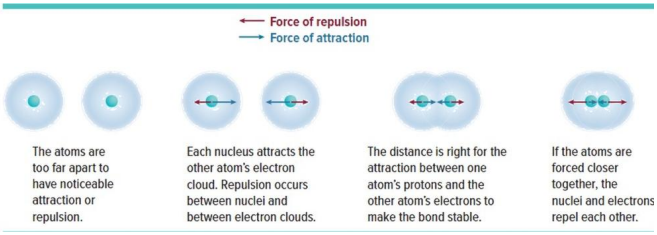
Question 11 [Practice]

A covalent bond results from:

A. transfer of electrons
B. attraction of cations to an electron sea
C. attraction between ions
D. sharing valence electrons

Question 12 [Practice]

According to Figure 2, a covalent bond forms when:



A. repulsion is at its maximum
B. the nuclei touch
C. the net attraction is at its maximum
D. the atoms are very far apart

Question 13 [Practice]

If two bonded atoms are forced closer than the bond length:

A. nothing happens
B. the bond becomes ionic
C. attraction increases
D. repulsion increases

Question 14 [Practice]

Which of the following is NOT a diatomic element?

A. He
B. O_2
C. Cl_2
D. N_2

Question 15 [Practice]

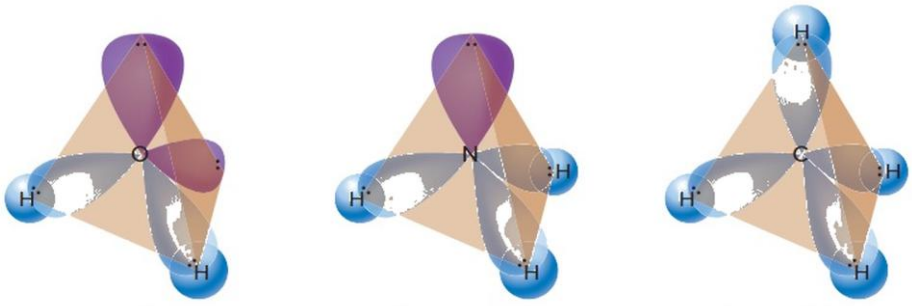
Covalent bonds most often form between:

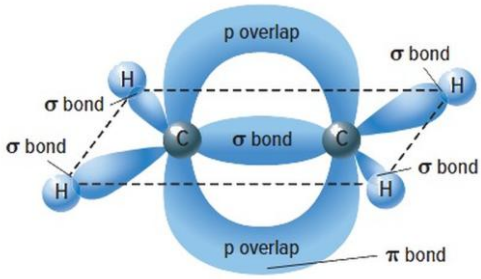
A. two nonmetals
B. a metal and a nonmetal
C. a metal and a noble gas
D. two metals

Q23

Learning outcome: Differentiate between sigma and pi bonds

Textbook: pp. 96–97 – Figures 7, 8, 9





Ethene

Key points from the lesson

- Sigma (σ) bond: the shared pair is centered between the two atoms; orbitals overlap end-to-end (s–s, s–p or p–p).
- Single covalent bonds are σ bonds: H_2O (2σ), NH_3 (3σ), CH_4 (4σ) – Figure 7.
- Pi (π) bond: parallel p orbitals overlap side-by-side; the shared pair is above and below the bond axis.
- Double bond = 1σ + 1π (O_2, C_2H_4). Triple bond = 1σ + 2π (N_2) – Figure 8.
- Ethene C_2H_4 (Figure 9): 4 C–H σ + 1 C–C σ + 1 π = 5σ and 1π.

Question 1 [Practice]

How many sigma bonds are in a methane molecule, CH_4?

A. 1
B. 2
C. 3
D. 4

Question 2 [Practice]

How many sigma and pi bonds are in the oxygen molecule, O_2?

A. 1 σ and 1 π
B. 2 σ and 0 π
C. 0 σ and 2 π
D. 1 σ and 2 π

Question 3 [Practice]

The numbers of sigma bonds in H_2O and NH_3 are…

A. 1 and 1
B. 2 and 3
C. 2 and 4
D. 3 and 3

Question 4 [Practice]

How many sigma and pi bonds are in the nitrogen molecule, N_2?

A. 1 σ and 1 π
B. 3 σ and 0 π
C. 1 σ and 2 π
D. 2 σ and 1 π

Question 5 [Practice]

Where is the shared electron pair of a pi bond?

A. Directly between the two nuclei
B. Above and below the line joining the two atoms
C. Inside one nucleus
D. Far away from both atoms

Question 6 [Practice]

Which statement correctly compares sigma and pi bonds?

A. A sigma bond forms by head-to-head overlap and is stronger than a pi bond.
B. A sigma bond forms by side-by-side overlap and is weaker than a pi bond.
C. A pi bond forms by head-to-head overlap and is stronger than a sigma bond.
D. Sigma and pi bonds form in the same way and have the same strength.

Question 7 [Practice]

A single covalent bond is…

A. always one sigma bond
B. always one pi bond
C. one sigma + one pi
D. two pi bonds

Question 8 [Practice]

How many sigma and pi bonds are present in an oxygen molecule, O_2?

A. 1 sigma and 1 pi
B. 2 sigma and 0 pi
C. 1 sigma and 2 pi
D. 0 sigma and 2 pi

Question 9 [Practice]

$COCl_2$ has one C=O bond and two C–Cl bonds. It contains…

A. 3 σ and 1 π
B. 2 σ and 2 π
C. 4 σ and 0 π
D. 3 σ and 2 π

Question 10 [Practice]

How many sigma and pi bonds are present in a nitrogen molecule, N_2?

A. 1 sigma and 2 pi
B. 3 sigma and 0 pi
C. 2 sigma and 1 pi
D. 1 sigma and 1 pi

Question 11 [Practice]

A sigma bond forms when orbitals overlap:

A. side-by-side
B. end-to-end
C. above the axis only
D. not at all

Question 12 [Practice]

A triple covalent bond consists of:

A. 3σ
B. 2σ + 1π
C. 1σ + 2π
D. 3π

Question 13 [Practice]

How many sigma bonds are in a methane (CH_4) molecule?

A. 1
B. 2
C. 3
D. 4

Question 14 [Practice]

How many σ and π bonds are in ethene, C_2H_4?

A. 4σ, 2π
B. 5σ, 1π
C. 6σ, 0π
D. 4σ, 1π

Question 15 [Practice]

A pi bond forms from:

A. parallel p orbitals overlapping side-by-side
B. s and p orbitals end-to-end
C. two s orbitals
D. a d orbital only

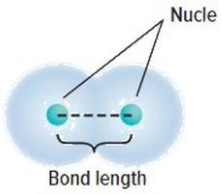
Q24

Learning outcome: Identify the relationship between the type of covalent bond (single, double, triple) and bond length, bond strength and dissociation energy

Textbook: p. 98 – Table 1

Table 1 Covalent Bond Type, Bond Length, and Bond-Dissociation Energy

Molecule	Bond Type	Bond Length	Bond-Dissociation Energy
F_2	single covalent	1.43×10^{-10} m	159 kJ/mol
O_2	double covalent	1.21×10^{-10} m	498 kJ/mol
N_2	triple covalent	1.10×10^{-10} m	945 kJ/mol



Key points from the lesson

- Bond length: the distance between two bonded nuclei at maximum attraction. It depends on atom size and the number of shared pairs.
- More shared pairs → shorter bond: F_2 (single) 1.43×10^{-10} m > O_2 (double) 1.21×10^{-10} m > N_2 (triple) 1.10×10^{-10} m.
- The shorter the bond, the stronger it is: single < double < triple.
- Bond-dissociation energy: energy required to break a specific covalent bond (always positive): F_2 159, O_2 498, N_2 945 kJ/mol.
- Breaking a bond absorbs energy; forming it releases the same amount.

Question 1 [Past exam 2025–26]

The table gives the bond-dissociation energies of N_2 (945 kJ/mol), F_2 (159 kJ/mol) and O_2 (498 kJ/mol). What is the order of the molecules by bond length from shortest to longest?

The following table represents the bond dissociation energy of a number of molecules.
What is the ascending order of the molecules in the table below according to bond length from **shortest** to **longest**?

الجدول الآتي يمثل طاقة تفكك الرابطة لعدد من الجزيئات.
ما الترتيب التصاعدي للجزيئات في الجدول أدناه حسب طول الرابطة من **الأقصر** إلى **الأطول**؟

molecule الجزيء	Bond Dissociation Energy طاقة تفكك الرابطة
N_2	945 kJ/mol
F_2	159 kJ/mol
O_2	498 kJ/mol

A. $N_2 \rightarrow O_2 \rightarrow F_2$
B. $O_2 \rightarrow N_2 \rightarrow F_2$
C. $O_2 \rightarrow F_2 \rightarrow N_2$
D. $F_2 \rightarrow O_2 \rightarrow N_2$

Question 2 [Practice]

Atoms that share two pairs of electrons form…

A. a single covalent bond
B. a double covalent bond
C. a triple covalent bond
D. an ionic bond

Question 3 [Practice]

What is the relationship between the length of a bond and its strength?

A. The shorter the bond, the stronger it is
B. The longer the bond, the stronger it is
C. There is no relationship
D. Length and strength are always equal

Question 4 [Practice]

Atoms that share three pairs of electrons form…

A. a single covalent bond
B. a double covalent bond
C. a triple covalent bond
D. a lone pair

Question 5 [Practice]

In general, the number of valence electrons an atom needs to form an octet equals…

A. the number of covalent bonds it can form
B. its number of protons
C. its number of lone pairs
D. its period number

Question 6 [Practice]

The bond in a nitrogen molecule, N_2, is…

A. single
B. double
C. triple
D. ionic

Question 7 [Practice]

Which description of the Lewis structure of CO_2 is correct?

A. C is central with two C=O double bonds; each O has two lone pairs.
B. C is central with two C–O single bonds; each O has three lone pairs.
C. O is central with one double bond to C and one to O.
D. C is central with one triple bond and one single bond; C has a lone pair.

Question 8 [Practice]

Which atoms often form multiple bonds with other nonmetals?

A. C, N, O and S
B. Na, K and Ca
C. He, Ne and Ar
D. Fe, Cu and Zn

Question 9 [Practice]

Book Q54: Can a binary compound with only sigma (single) bonds display resonance?

A. Only with lone pairs
B. No — resonance needs both single and double bonds
C. Yes, always
D. Only if it is an ion

Question 10 [Practice]

Which elements often form double and triple bonds?

A. C, N, O and S
B. H and F
C. Na and Cl
D. He and Ne

Question 11 [Practice]

Which molecule has the shortest bond?

A. All are equal
B. N_2
C. O_2
D. F_2

Question 12 [Practice]

Using Table 1, what is the bond-dissociation energy of O_2?

Table 1 Covalent Bond Type, Bond Length, and Bond-Dissociation Energy

Molecule	Bond Type	Bond Length	Bond-Dissociation Energy
F_2	single covalent	1.43×10^{-10} m	159 kJ/mol
O_2	double covalent	1.21×10^{-10} m	498 kJ/mol
N_2	triple covalent	1.10×10^{-10} m	945 kJ/mol

A. 159 kJ/mol
B. 498 kJ/mol
C. 945 kJ/mol
D. 1.21 kJ/mol

Question 13 [Practice]

Which relationship is correct?

A. Shorter bond → weaker bond
B. Shorter bond → greater bond-dissociation energy
C. Single bonds are the strongest
D. Bond length does not affect strength

Question 14 [Practice]

Breaking a covalent bond:

A. releases energy
B. absorbs energy
C. involves no energy
D. creates ions only

Question 15 [Practice]

Which order of bond strength is correct?

A. $F_2 < N_2 < O_2$
B. $N_2 < O_2 < F_2$
C. $O_2 < F_2 < N_2$
D. $F_2 < O_2 < N_2$

Q25

Learning outcome: Name a binary molecular compound from its molecular formula

Textbook: pp. 100–101 – Table 2 + Example 2

Table 2 Prefixes in Covalent Compounds

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Key points from the lesson

- A binary molecular compound contains only two nonmetals (no metal, no ions).
- Rule 1: the first element is named with its full name. Rule 2: the second element = root + -ide.
- Rule 3: prefixes give the number of atoms: mono 1, di 2, tri 3, tetra 4, penta 5, hexa 6, hepta 7, octa 8, nona 9, deca 10.
- The first element never uses mono- (CO = carbon monoxide). Drop a vowel when two vowels meet (monoxide, not monooxide).
- Example 2: P_2O_5 = diphosphorus pentoxide; N_2O = dinitrogen monoxide.

Question 1 [Past exam 2025–26]

What is the name of the covalent compound with the formula N_2O_4?

A. Dinitrogen tetroxide
B. Tetranitrogen dioxide
C. Nitrogen dioxide
D. Nitrogen tetroxide

Question 2 [Practice]

Formula of chlorine trifluoride? (Book Q27)

A. Cl_3F B. ClF_3 C. ClF D. Cl_2F_3

Question 3 [Practice]

When writing a formula from the name of a binary molecular compound, the prefixes give…

A. the subscripts
B. the charges
C. the number of bonds
D. the mass

Question 4 [Practice]

Formula of diphosphorus trioxide? (Book Q28)

A. P_3O_2 B. PO_3 C. P_2O_5 D. P_2O_3

Question 5 [Practice]

Book Q23 (Challenge): general formula for a compound of a group 2 element (M) and a group 17 element (X)

A. MX B. M_2X C. MX_2 D. M_2X_3

Question 6 [Practice]

Formula of disulfur decafluoride? (Book Q29)

A. S_2F_{10} B. $S_{10}F_2$ C. SF_6 D. S_2F_8

Question 7 [Practice]

What is the formula of the compound formed from lithium and oxygen (Q62)?

A. LiO
B. Li_2O
C. LiO_2
D. Li_2O_2

Question 8 [Practice]

How would you name N_2O_4? (Book Q34)

A. dinitrogen tetroxide
B. nitrogen tetroxide
C. dinitrogen oxide
D. tetranitrogen dioxide

Question 9 [Practice]

When you write a formula from the name of a binary molecular compound, the prefixes tell you…

A. the charges of the ions
B. the subscripts (number of each atom)
C. which element is the acid
D. the number of bonds

Question 10 [Practice]

Name CO_2. (Book Q14)

A. carbon dioxide
B. monocarbon dioxide
C. carbon oxide
D. dicarbon oxide

Question 11 [Practice]

What is the name of CO_2?

A. carbon(II) oxide
B. carbon dioxide
C. carbon oxide
D. monocarbon dioxide

Question 12 [Practice]

What is the name of NF_3?

A. nitrogen fluoride
B. nitrogen trifluoride
C. mononitrogen trifluorine
D. trinitrogen fluoride

Question 13 [Practice]

What is the name of P_2O_5?

A. phosphorus oxide
B. pentaphosphorus dioxide
C. diphosphorus pentaoxide
D. diphosphorus pentoxide

Question 14 [Practice]

Which prefix means 7?

A. hepta-
B. hexa-
C. octa-
D. nona-

Question 15 [Practice]

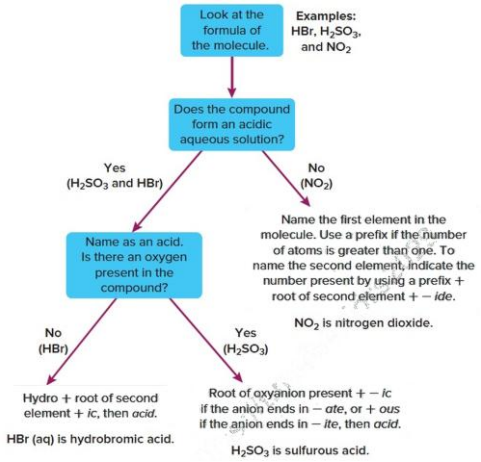
What is the name of CCl_4?

A. carbon tetrachloride
B. carbon chloride
C. carbon(IV) chloride
D. tetracarbon chloride

Q26

Learning outcome: Determine the chemical formula of a compound from its name

Textbook: pp. 102–104 – Table 3 + Figure 12



Key points from the lesson

- An acid is a compound that produces hydrogen ions (H^+) in water solution.
- Binary acid (hydrogen + one other element): hydro- + root + -ic, then "acid": HCl → hydrochloric acid, HBr → hydrobromic acid. HCN → hydrocyanic acid.
- Oxyacid (hydrogen + oxyanion): -ate → -ic; -ite → -ous. The hydrogen is not part of the name.
- Examples: HNO_3 nitric acid, HNO_2 nitrous acid, $HClO_3$ chloric acid, $HClO_2$ chlorous acid, H_2SO_4 sulfuric acid, H_2SO_3 sulfurous acid.
- Figure 12: not acidic → name with prefixes (NO_2 = nitrogen dioxide); acidic → check for oxygen.

Question 1 [Practice]

What is the formula for diarsenic trioxide? (Book Q18)

A. As_2O_3
B. As_3O_2
C. AsO_3
D. As_2O

Question 2 [Practice]

Write the formula: tetraphosphorus decoxide.

A. P_4O_{10}
B. P_2O_5
C. P_4O_9
D. $P_{10}O_4$

Question 3 [Practice]

Write the formula: dinitrogen trioxide.

A. N_2O_3
B. N_3O_2
C. NO_3
D. N_2O

Question 4 [Practice]

Why is the overall charge of a formula unit zero?

A. It represents the whole crystal, which is electrically neutral
B. Ions lose their charges in a crystal
C. Only cations are counted
D. Electrons are shared

Question 5 [Practice]

What is the formula for the iodide ion?

A. I^+
B. I^{2-}
C. I^-
D. I^{3-}

Question 6 [Practice]

What is the name of HI dissolved in water?

A. iodous acid
B. hydrogen iodate
C. hydroiodic acid
D. iodic acid

Question 7 [Practice]

What is the name of H_2SO_4?

A. sulfurous acid
B. hydrosulfuric acid
C. sulfuric acid
D. hydrogen sulfite

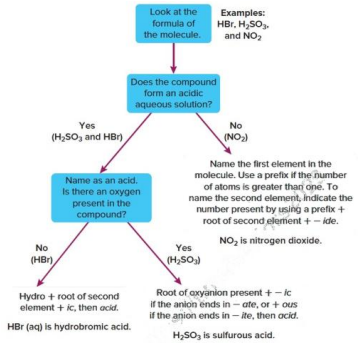
Question 8 [Practice]

What is the name of $HClO_2$?

A. chloric acid
B. perchloric acid
C. hydrochloric acid
D. chlorous acid

Question 9 [Practice]

Using Figure 12, H_2SO_3 is named:



A. sulfuric acid
B. hydrosulfuric acid
C. sulfurous acid
D. sulfur trioxide

Question 10 [Practice]

What is the formula of nitrous acid?

A. HN
B. HNO_2
C. H_2NO_2
D. HNO_3

Q27

Learning outcome: Determine the chemical formula of a compound from its name

Textbook: p. 103 – Table 4

Table 4 Formulas and Names of Some Covalent Compounds

Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

Key points from the lesson

- The name of a molecular compound reveals its composition; the prefixes give the subscripts in the formula.
- Examples: dinitrogen tetrahydride → N_2H_4; diarsenic trioxide → As_2O_3; sulfur hexafluoride → SF_6.
- No prefix on the first element means one atom (carbon monoxide → CO).
- Table 4 – common names: water H_2O (dihydrogen monoxide), ammonia NH_3 (nitrogen trihydride), hydrazine N_2H_4, muriatic acid HCl (hydrochloric acid), aspirin $C_9H_8O_4$.
- Acid names give formulas too: periodic acid → HIO_4 (from periodate IO_4^-).

Question 1 [Practice]

The formula unit of magnesium chloride is $MgCl_2$ because the ions exist in a ratio of…

A. 1 : 1
B. 2 : 1
C. 1 : 2
D. 1 : 3

Question 2 [Practice]

Write the formula: sulfur hexafluoride.

A. SF_6
B. S_6F
C. SF_4
D. S_2F_6

Question 3 [Practice]

Which cation would form a stronger ionic bond with Cl^-?

A. Na^+ B. Ca^{2+}

Question 4 [Practice]

What is the formula for the beryllium ion?

A. Be^+
B. Be^{2+}
C. Be^{2-}
D. Be^{3+}

Question 5 [Practice]

What is the formula for the nitride ion?

A. N^{3-}
B. N^{3+}
C. N^{2-}
D. N^-

Question 6 [Practice]

What is the formula of dinitrogen tetrahydride?

A. N_2H_4
B. N_4H_2
C. N_2H_2
D. NH_4

Question 7 [Practice]

What is the formula of diarsenic trioxide?

A. AsO_3
B. As_2O_3
C. As_3O_2
D. As_2O

Question 8 [Practice]

According to Table 4, the common name of NH_3 is:

Table 4 Formulas and Names of Some Covalent Compounds

Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

A. ammonia
B. hydrazine
C. muriatic acid
D. water

Question 9 [Practice]

What is the molecular compound name of water?

A. hydrogen hydroxide
B. dihydrogen dioxide
C. dihydrogen monoxide
D. hydrogen oxide

Question 10 [Practice]

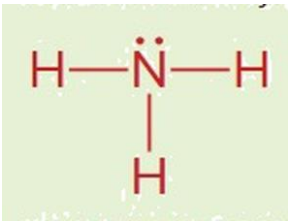
What is the formula of sulfur hexafluoride?

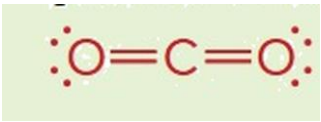
A. SF_6
B. S_6F
C. S_2F_6
D. SF_4

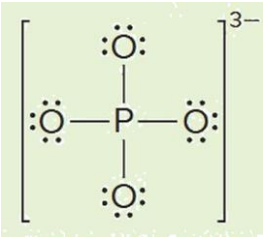
Q28

Learning outcome: Draw Lewis structures for covalent compounds with single and multiple bonds

Textbook: pp. 107–110 – Examples 3, 4, 5







Key points from the lesson

- 1) Central atom: the one with the least attraction for shared electrons (H is always terminal). 2) Count all valence electrons.
- 3) Bonding pairs = electrons ÷ 2. 4) Put a single bond between the central atom and each terminal atom.
- 5) Place remaining pairs as lone pairs (terminal atoms first, then central). 6) If the central atom lacks an octet, turn lone pairs into double or triple bonds.
- NH_3: 8 e^- → 3 N–H bonds + 1 lone pair on N. CO_2: 16 e^- → O=C=O (two double bonds).
- Polyatomic ion: add the negative charge (subtract a positive charge). PO_4^{3-}: 5 + 4(6) + 3 = 32 e^- = 16 pairs; draw brackets and the charge.

Question 1 [Past exam 2025–26]

Which electron configuration represents an element that forms three covalent bonds with other nonmetal atoms?

A. $1s^2\ 2s^2\ 2p^3$
B. $1s^2\ 2s^2\ 2p^6\ 3s^1$
C. $1s^2\ 2s^2\ 2p^5$
D. $1s^2\ 2s^2\ 2p^4$

Question 2 [Practice]

In the Lewis structure of PH_3, the phosphorus atom has…

A. 3 bonding pairs and 1 lone pair
B. 4 bonding pairs
C. 3 bonding pairs and no lone pairs
D. 2 bonding pairs and 2 lone pairs

Question 3 [Practice]

In CO_2, after placing the single bonds and the lone pairs on O, carbon has only 4 electrons. What do you do?

A. Convert a lone pair from each O into a double bond
B. Add more electrons
C. Remove an oxygen
D. Put a lone pair on carbon

Question 4 [Practice]

How many covalent bonds can phosphorus (group 15) usually form?

A. 1
B. 3
C. 5
D. 15

Question 5 [Practice]

A group 1 element (X) bonds with a group 16 element (Y). The general Lewis structure is…

A. X–Y with 3 lone pairs on Y
B. Y–X–Y
C. X=Y
D. X–Y–X with 2 lone pairs on Y

Question 6 [Practice]

How many valence electrons are available for bonding in NF_3?

A. 8
B. 24
C. 26
D. 28

Question 7 [Practice]

In the octet Lewis structure of PO_4^{3-} (four P–O single bonds), how many lone pairs are there in total?

A. 12
B. 8
C. 16
D. 4

Question 8 [Practice]

How many bonding pairs (total pairs) are available in CO_2?

A. 4
B. 6
C. 8
D. 16

Question 9 [Practice]

How many single covalent bonds does an atom of a group 15 element usually form?

A. 1
B. 4
C. 3
D. 5

Question 10 [Practice]

How many bonds does krypton (Kr) usually form?

A. 0
B. 1
C. 2
D. 8

Question 11 [Practice]

How many valence electrons are available for bonding in CO_2?

A. 12
B. 14
C. 16
D. 18

Question 12 [Practice]

How many lone pairs are on the nitrogen atom in NH_3?

A. 0
B. 1
C. 2
D. 3

Question 13 [Practice]

In the Lewis structure of CO_2, each C–O bond is a:

A. single bond
B. double bond
C. triple bond
D. ionic bond

Question 14 [Practice]

How many valence electrons are in the phosphate ion, PO_4^{3-}?

A. 29
B. 32
C. 35
D. 26

Question 15 [Practice]

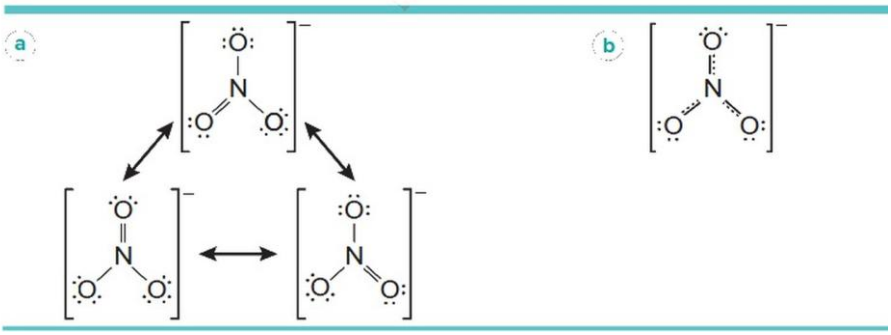
Which atom is the central atom in NH_4^+, and how many valence electrons does the ion have?

A. N, 10
B. H, 9
C. N, 8
D. N, 9

Q29

Learning outcome: Identify resonance structures (O3, NO3−, NO2−, SO3 2−, CO3 2−)

Textbook: pp. 111–112 – Figure 14



Key points from the lesson

- Resonance occurs when more than one valid Lewis structure can be written for a molecule or ion.
- Resonance structures differ only in the position of electron pairs, never in the positions of the atoms.
- NO_3^- has three equivalent structures (Figure 14a). The actual ion is an average of them (Figure 14b).
- Measured bond lengths are identical: shorter than single bonds but longer than double bonds.
- Examples that show resonance: O_3, NO_3^-, NO_2^-, SO_3^{2-}, CO_3^{2-}.

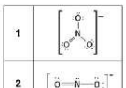
Question 1 [Past exam 2025–26]

Which of the following is NOT a resonance structure for the nitrate ion NO_3^-?

Which of the following is not a resonance structure for the nitrate ion NO_3^-?

أي مما يلي لا يعتبر من تراكيب الرنين لأيون النيترات NO_3^-؟

العدد الذري Atomic number	العنصر Element
8	Oxygen (O) الأكسجين
7	Nitrogen (N) النيتروجين

1	
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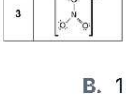
3	

A. 2 only
B. 1 only
C. 1 and 2
D. 2 and 3

Question 2 [Practice]

How many valid resonance structures can be drawn for the nitrate ion, NO_3^-?

A. 2
B. 1
C. 4
D. 3

Question 3 [Practice]

In the carbonate ion, CO_3^{2-}, the three C–O bonds are…

A. identical, between a single and a double bond in length
B. three double bonds
C. one double and two single bonds of different lengths
D. three single bonds

Question 4 [Practice]

The sulfite ion, SO_3^{2-}, drawn with one S=O double bond, has how many resonance structures?

A. 4
B. 1
C. 2
D. 3

Question 5 [Practice]

Ozone, O_3, has how many resonance structures?

A. 4
B. 2
C. 1
D. 3

Question 6 [Practice]

Which species shows resonance?

A. NH_3
B. NO_2^-
C. H_2O
D. CH_4

Question 7 [Practice]

Which statement about resonance structures is correct?

A. The molecule flips quickly from one structure to another
B. Only the positions of electron pairs differ; the atoms stay in place
C. The atoms change positions between structures
D. Each structure has a different number of electrons

Question 8 [Practice]

Which of these does NOT exhibit resonance?

A. CO_3^{2-}
B. CH_4
C. O_3
D. NO_2^-

Question 9 [Practice]

How many resonance structures does the nitrite ion, NO_2^-, have?

A. 4
B. 2
C. 3
D. 1

Question 10 [Practice]

Which group lists only species that show resonance?

A. HCl, NO_2^-, H_2
B. O_3, NO_3^-, CO_3^{2-}
C. O_3, H_2O, CO_3^{2-}
D. NH_3, NO_3^-, CH_4

Question 11 [Practice]

Resonance structures of a molecule differ in the positions of:

A. the nuclei
B. the atoms
C. the protons
D. the electron pairs

Question 12 [Practice]

How many equivalent resonance structures does the nitrate ion (NO_3^-) have?

A. 1
B. 2
C. 3
D. 4

Question 13 [Practice]

Which molecule exhibits resonance?

A. CH_4
B. O_3
C. H_2O
D. NH_3

Question 14 [Practice]

The actual N–O bonds in NO_3^- are:

A. all longer than single bonds
B. all identical, between single and double in length
C. one double and two single
D. all triple

Question 15 [Practice]

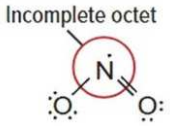
How many resonance structures can be drawn for the nitrite ion, NO_2^-?

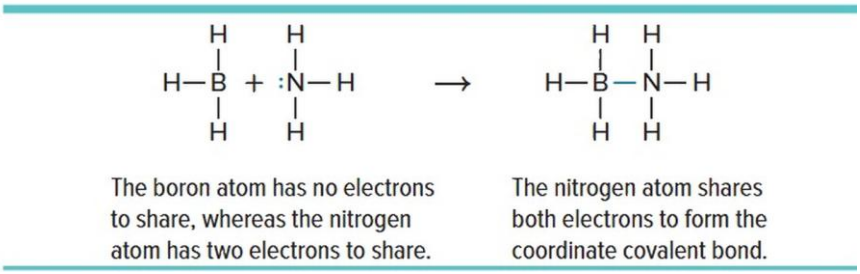
A. 1
B. 2
C. 3
D. 4

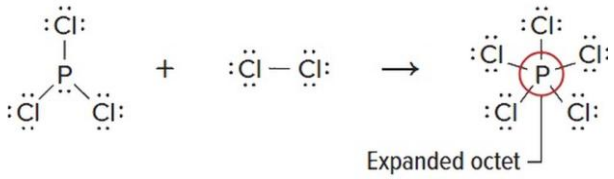
Q30

Learning outcome: Determine exceptions to the octet rule (odd number of valence electrons, sub-octets, coordinate covalent bonds, expanded octets)

Textbook: pp. 112–114 – Figures 16, 17 + Example 6







Key points from the lesson

- Odd number of valence electrons: NO_2 (5 + 12 = 17 e^-), ClO_2 and NO cannot form an octet around every atom (Figure 15).
- Suboctet: fewer than eight electrons around an atom. In BH_3, boron has only 6 electrons; such compounds are reactive.
- Coordinate covalent bond: one atom donates both shared electrons. NH_3 donates its lone pair to BH_3 (Figure 16).
- Expanded octet: central atom in period 3 or higher uses d orbitals to hold more than 8 electrons: PCl_5 (10 e^-), SF_6 (12 e^-).
- Example 6: XeF_4 has 36 e^-; Xe gets 4 bonds + 2 lone pairs = 12 electrons.

Question 1 [Practice]

ClO_2 has 19 valence electrons. Which exception to the octet rule does it show?

A. odd number of valence electrons
B. expanded octet
C. suboctet
D. coordinate covalent bond

Question 2 [Practice]

In PCl_5, phosphorus forms five bonds. Which exception to the octet rule does it show?

A. expanded octet
B. suboctet
C. odd number of valence electrons
D. it obeys the octet rule

Question 3 [Practice]

In BH_3 (and BF_3), boron forms only three bonds. Which exception to the octet rule does it show?

A. suboctet (fewer than eight electrons)
B. expanded octet
C. odd number of valence electrons
D. it obeys the octet rule

Question 4 [Practice]

NO_2 has 17 valence electrons. Which exception to the octet rule does it show?

A. odd number of valence electrons
B. expanded octet
C. suboctet
D. it obeys the octet rule

Question 5 [Practice]

NO has 11 valence electrons. Which exception to the octet rule does it show?

A. odd number of valence electrons
B. expanded octet
C. suboctet
D. it obeys the octet rule

Question 6 [Practice]

PF_5 is an example of…

A. an expanded octet
B. a suboctet
C. an odd number of electrons
D. a coordinate covalent bond

Question 7 [Practice]

In ethylene, C_2H_4 (12 e^-, 6 pairs), the bond between the two carbon atoms is…

A. single
B. double
C. triple
D. ionic

Question 8 [Practice]

In $BH_3 + NH_3 \rightarrow H_3BNH_3$, which atom donates the electron pair?

A. H
B. No atom
C. N
D. B

Question 9 [Practice]

In ball-and-stick and space-filling models, atoms of each element are identified by…

A. a representative colour
B. their mass
C. a number
D. their charge

Question 10 [Practice]

An expanded octet means the central atom has…

A. fewer than eight electrons
B. an odd number of electrons
C. exactly eight electrons
D. more than eight valence electrons

Question 11 [Practice]

NO_2 does not obey the octet rule because:

A. it is ionic
B. O has 10 electrons
C. N has an expanded octet
D. it has an odd number of valence electrons

Question 12 [Practice]

How many electrons surround boron in BH_3?

A. 4
B. 6
C. 8
D. 10

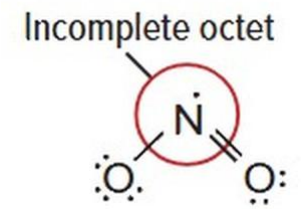
Question 13 [Practice]

In a coordinate covalent bond:

A. electrons are transferred completely
B. one atom donates both shared electrons
C. no electrons are shared
D. each atom gives one electron

Question 14 [Practice]

How many electrons surround P in PCl_5 (Figure 17)?



A. 6
B. 8
C. 10
D. 12

Question 15 [Practice]

Sulfur can form SF_6, but oxygen cannot form OF_6, because:

A. S has fewer electrons
B. O is in group 18
C. O is a metal
D. S is in period 3 and has d orbitals available

**********the end 😊