الملف الخطة الأسبوعية للأسبوع الخامس الحلقة الثانية في مدرسة أبو أيوب الأنصاري
موقع المناهج ⇦ المناهج الإماراتية ⇦ ملفات مدرسية ⇦ المدارس ⇦ الفصل الأول

Department of Education and Knowledge

Al Dahmaa School

Chemistry grade 10 advanced

هيكلة اختبار الكيمياء عاشر متقدم – الفصل

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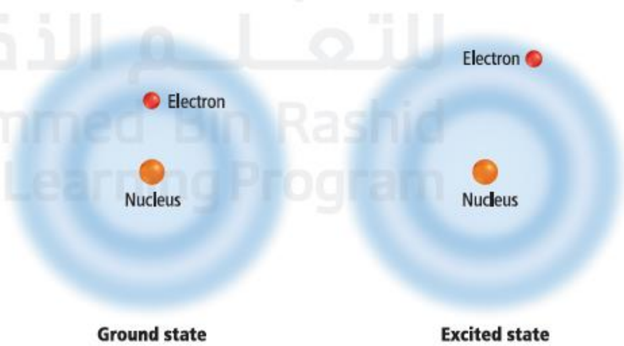
ال	ناتج التعلم***	Example/Exercise مثال/تمرين	Page الصفحة
1	Differentiate between the ground and excited states of an atom تقارن بين الحالة الأرضية وحالة الاستثارة للذرة	Figure 10 , Figure 11 الشكل 10 والشكل 11	14 - 15

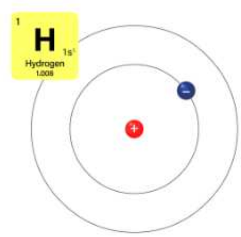
Figure 10 The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.



Fill in the blank with the correct term from the Word Bank.

Word Bank			
excited	stable	ionized	gaseous

A hydrogen atom is stable at the ground state.



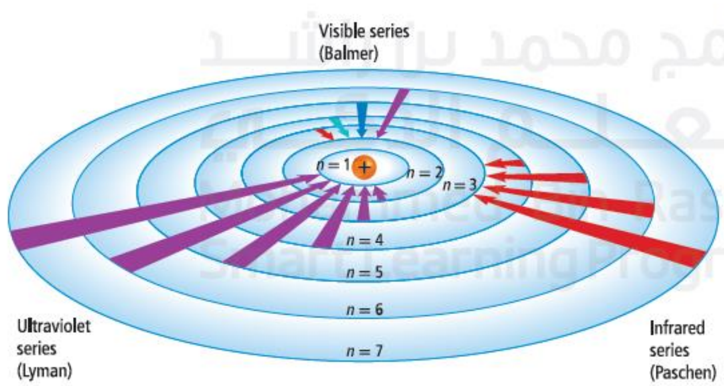


Figure 11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

- Visible series (Balmer) from n=6 to n= 2
- Ultraviolet series (Lyman) from n=7 to n= 1
- Infrared series (Paschen) from n=7 to n=3

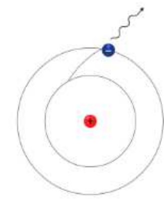
Select your answer from the drop-down menu.

The Lyman series refers to emissions that fit into the ultraviolet range.

Select the correct answer.

Study the diagram showing an atom alongside.

In which state is this atom?

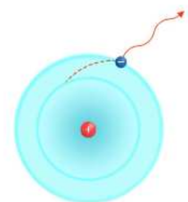


- neutral
- excited
- absorption
- ground

Select the correct answer.

Study the diagram showing an atom alongside.

When an atom drops from its excited state, it emits a(n) _____.

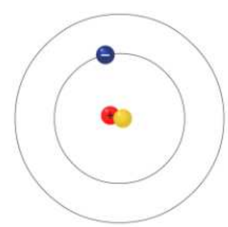


- proton
- photon
- electron
- neutron

Select the correct answer.

Study the diagram of a hydrogen atom (H) alongside.

Which state does this diagram show?



- neutral
- ground
- excited
- absorption

Select the correct answer.

Which series refers to emissions in the infrared range?

- Atomic Emission Spectrum
- Paschen series
- Balmer series
- Lyman series

Drag and drop the correct terms.

The Balmer series refers to emissions that fit into the visible range.

These emissions occur when electrons drop from a larger orbit back to the $n =$ 2 level.

Lyman | Paschen | 3 | 1

2	Explain the aufbau principle , the Pauli exclusion principle,and Hund's rule	Text Book	24 - 25 - 26
	يفسر مبدأ أوفباو ومبدأ باولي للاستبعاد وقاعدة هوند	نص الكتاب	

Ground-State Electron Configuration

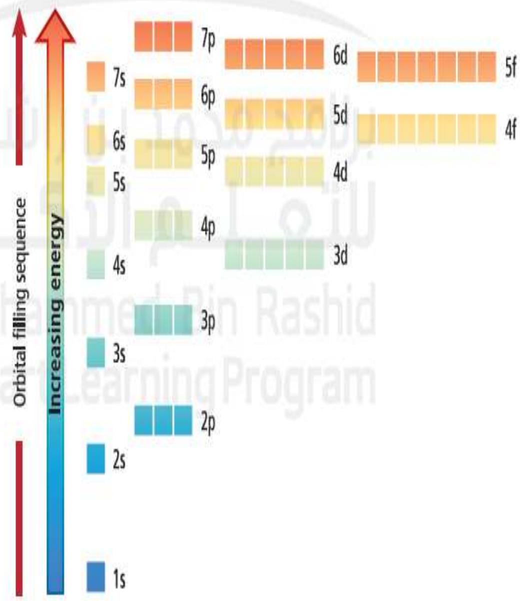
When you consider that atoms of the heaviest elements contain more than 100 electrons, the idea of determining electron arrangements in atoms with many electrons seems daunting. Fortunately, all atoms can be described with orbitals similar to hydrogen's. This allows us to describe arrangements of electrons in atoms using a few specific rules.

The arrangement of electrons in an atom is called the atom's **electron configuration.** Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest energy possible. The most stable, lowest-energy arrangement of the electrons is called the element's ground-state electron configuration. Three rules, or principles–the aufbau principle, the Pauli exclusion principle, and Hund's rule–define how electrons can be arranged in an atom's orbitals.

The aufbau principle The **aufbau principle** states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element's ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in **Figure 18.** In the diagram, each box represents an atomic orbital.

Figure 18 The aufbau diagram shows the energy of each sublevel relative to the energy of other sublevels. Each box on the diagram represents an atomic orbital.

Determine *Which sublevel has the greater energy, 4d or 5p?*



The Pauli exclusion principle Electrons in orbitals can be represented by arrows in boxes. Each electron has an associated spin, similar to the way a top spins on its point. Like the top, the electron is able to spin in only one of two directions. An arrow pointing up $\boxed{\uparrow}$ represents the electron spinning in one direction, and an arrow pointing down $\boxed{\downarrow}$ represents the electron spinning in the opposite direction. An empty box $\boxed{}$ represents an unoccupied orbital, a box containing a single up arrow $\boxed{\uparrow}$ represents an orbital with one electron, and a box containing both up and down arrows $\boxed{\uparrow\downarrow}$ represents a filled orbital.

The **Pauli exclusion principle** states that a maximum of two electrons can occupy a single atomic orbital, but only if the electrons have opposite spins. Austrian physicist Wolfgang Pauli (1900-1958) proposed this principle after observing atoms in excited states. An atomic orbital containing paired electrons with opposite spins is written as $\boxed{\uparrow\downarrow}$. Because each orbital can contain, at most, two electrons, the maximum number of electrons related to each principal energy level equals $2n^2$.

Hund's rule The fact that negatively charged electrons repel each other has an important impact on the distribution of electrons in equal-energy orbitals. **Hund's rule** states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. For example, let the boxes below represent the 2p orbitals. One electron enters each of the three 2p orbitals before a second electron enters any of the orbitals. The sequence in which six electrons occupy three p orbitals is shown below.

1. $\boxed{\uparrow}\boxed{}\boxed{}$ **2.** $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$ **3.** $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

4. $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$ **5.** $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$ **6.** $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$

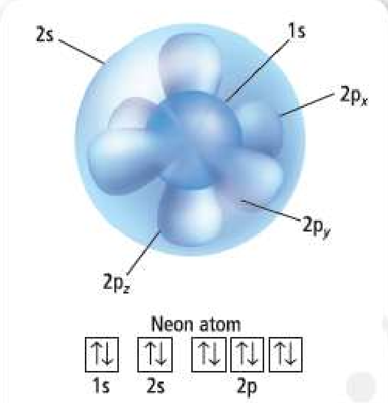
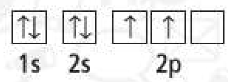


Figure 19 The 1s, 2s, and 2p orbitals of a neon atom overlap.
Determine how many electrons a neon atom has.

Electron Arrangement

You can represent an atom's electron configuration using one of two convenient methods: orbital diagrams or electron configuration notation.

Orbital diagrams As mentioned earlier, electrons in orbitals can be represented by arrows in boxes. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, shown below, contains two electrons in the 1s orbital, two electrons in the 2s orbital, and one electron in two of three separate 2p orbitals. The third 2p orbital remains unoccupied.



Electron configuration notation The electron configuration notation designates the principal energy level and energy sublevel associated with each of the atom's orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written $1s^2 2s^2 2p^2$. Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 4. Figure 19** illustrates how the 1s, 2s, $2p_x$, $2p_y$, and $2p_z$ orbitals illustrated previously in **Figure 17** overlap in the neon atom.

المطلوب من ناتج التعلم الثاني : تعريف مبدأ باولي واوفباو وقاعدة هوند للتوزيع الالكتروني (فيما يلي بعض الأسئلة للناتج التعلم الثاني)

Select the best answer.

The sequence of filling the energy sublevels within the third principal energy level is: _____

- 1s 2s 2p 3s 3p 3d
- 3d 3p 3s 2p 2s 1s
- 3s 3p 3d ✓
- 3d 3p 3s

Rule or Principle for Electron Configuration	Description
aufbau principle	electrons fill lowest-energy orbitals before higher-energy orbitals
Pauli exclusion principle	two electrons can occupy the same atomic orbital if they spin in opposite directions
Hund's rule	single electrons fill orbitals of a sublevel first before pairing with an electron of opposite spin

Select the best answer.

The atomic number of the element oxygen (O) is 8. Identify the orbital diagram that represents the correct order in which the electrons fill up the orbitals.

- ○ 2p ↑↓ ↑↓ ↑↓
2s ↑
1s ↑
- ○ 2p ↑↓ ↑↓ ↑↓
2s ↓
1s ↓
- ○ 2p ↑↓ ↑↓ ☐
2s ↑↓
1s ↑↓
- ◉ 2p ↑↓ ↑ ↑
2s ↑↓
1s ↑↓ ✓

3	Use the electron configurations notation, orbital notation, and noble gas notation of an element (Z1-36) to identify the location of the element in the periodic table (period,group,block)	Example 1 , Applications 8 , 9 , 10	54
	يستخدم الترتيب الإلكتروني وترميز الفلك وترميز الغاز النبيل (Z1 -36) لتحديد موقع العنصر في الجدول الدوري(المجموعة – الدورة – المجمع)	مثال 1 وتطبيقات 8 و 9 و 10	

EXAMPLE 1

ELECTRON CONFIGURATION AND THE PERIODIC TABLE Strontium, which is used to produce red fireworks, has an electron configuration of [Kr]$5s^2$. Without using the periodic table, determine the group, period, and block of strontium.

1 ANALYZE THE PROBLEM

You are given the electron configuration of strontium.

Known	Unknown
Electron configuration = [Kr]$5s^2$	Group = ?
	Period = ?
	Block = ?

2 SOLVE FOR THE UNKNOWN

The s^2 indicates that strontium's valence electrons fill the s sublevel. Thus, strontium is in group 2 of the s-block.

For representative elements, the number of valence electrons can indicate the group number.

The 5 in $5s^2$ indicates that strontium is in period 5.

The number of the highest energy level indicates the period number.

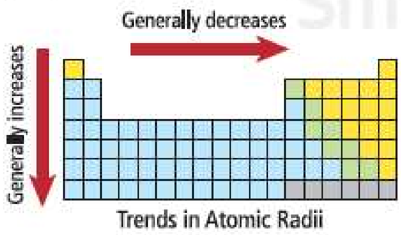
3 EVALUATE THE ANSWER

The relationships between electron configuration and position on the periodic table have been correctly applied.

APPLICATIONS

8. Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.
   a. [Ne]$3s^2$ b. [He]$2s^2$ c. [Kr]$5s^2 4d^{10} 5p^5$
9. What are the symbols for the elements with the following valence electron configurations?
   a. s^2d^1 b. s^2p^3 c. s^2p^6
10. **Challenge** Write the electron configuration of the following elements.
    a. the group 2 element in the fourth period
    b. the group 12 element in the fourth period
    c. the noble gas in the fifth period
    d. the group 16 element in the second period

4	Predict the trends of the properties of elements across a period and a group of the periodic table	Text Book	56- 57-58-59
	يتنبأ بدورية خواص العناصر في الدورة والمجموعة في الجدول الدوري	نص الكتاب	



■ **Figure 12** Atomic radii generally decrease from left to right in a period and generally increase as you move down a group.

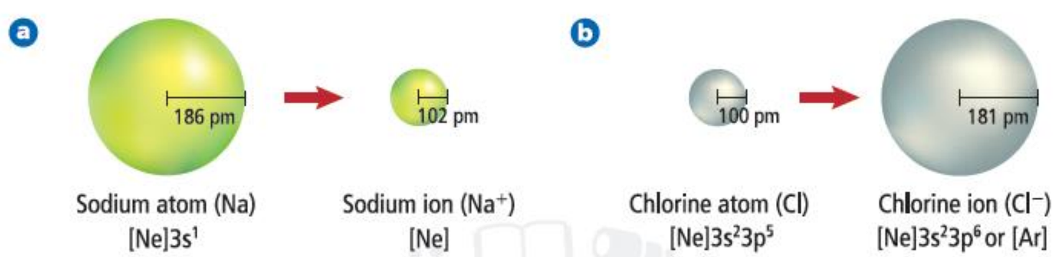
Trends within groups Atomic radii generally increase as you move down a group. The nuclear charge increases, and electrons are added to orbitals corresponding to successively higher principal energy levels. However, the increased nuclear charge does not pull the outer electrons toward the nucleus to make the atom smaller.

Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, the atom becomes larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the pull of the increased nuclear charge. Also, as additional orbitals between the nucleus and the outer electrons are occupied, these electrons shield the outer electrons from the nucleus. **Figure 12** summarizes the group and period trends.

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge. You will learn about ions later, but for now, consider how the formation of an ion affects the size of an atom.

When atoms lose electrons and form positively charged ions, they always become smaller. The reason is twofold. The electron lost from the atom will almost always be a valence electron. The loss of a valence electron can leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now-fewer number of remaining electrons decreases. As a result, they experience a greater nuclear charge allowing these remaining electrons to be pulled closer to the positively charged nucleus.



■ **Figure 13** The size of atoms varies greatly when they form ions.
a. Positive ions are smaller than the neutral atoms from which they form.
b. Negative ions are larger than the neutral atoms from which they form.

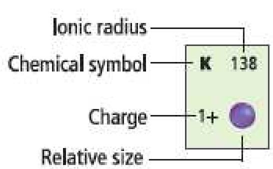
When atoms gain electrons and form negatively charged ions, they become larger. The addition of an electron to an atom increases the electrostatic repulsion between the atom's outer electrons, forcing them to move farther apart. The increased distance between the outer electrons results in a larger radius.

Figure 13a illustrates how the radius of sodium decreases when sodium atoms form positive ions, and **Figure 13b** shows how the radius of chlorine increases when chlorine atoms form negative ions.

Trends within periods The ionic radii of most of the representative elements are shown in **Figure 14**. Note that elements on the left side of the table form smaller positive ions, and elements on the right side of the table form larger negative ions. In general, as you move from left to right across a period, the size of the positive ions gradually decreases. Then, beginning in group 15 or 16, the size of the much-larger negative ions also gradually decreases.

Figure 14 The ionic radii of most of the representative elements are shown in picometers (10^{-12} m).

Explain *why the ionic radii increase for both positive and negative ions as you move down a group.*



Period	1	2	13	14	15	16	17
2	Li 76, 1+	Be 31, 2+	B 20, 3+	C 15, 4+	N 146, 3–	O 140, 2–	F 133, 1–
3	Na 102, 1+	Mg 72, 2+	Al 54, 3+	Si 41, 4+	P 212, 3–	S 184, 2–	Cl 181, 1–
4	K 138, 1+	Ca 100, 2+	Ga 62, 3+	Ge 53, 4+	As 222, 3–	Se 198, 2–	Br 196, 1–
5	Rb 152, 1+	Sr 118, 2+	In 81, 3+	Sn 71, 4+	Sb 62, 5+	Te 221, 2–	I 220, 1–
6	Cs 167, 1+	Ba 135, 2+	Tl 95, 3+	Pb 84, 4+	Bi 74, 5+		

Trends within groups As you move down a group, an ion's outer electrons are in orbitals corresponding to higher principal energy levels, resulting in a gradual increase in ionic size. Thus, the ionic radii of both positive and negative ions increase as you move down a group. The group and period trends in ionic radii are summarized in **Figure 15**.

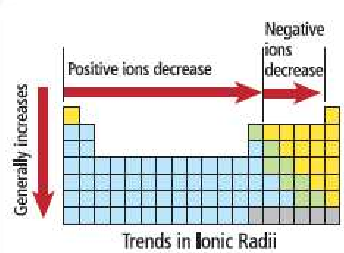


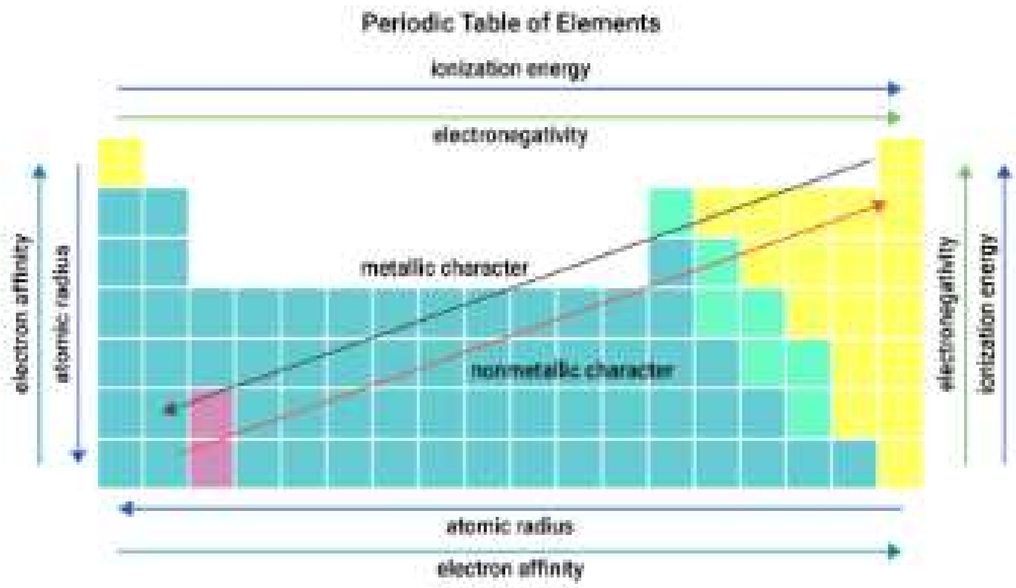
Figure 15 The diagram summarizes the general trends in ionic radii.

Ionization Energy

To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge of the nucleus and the negative charge of the electron. **Ionization energy** is defined as the energy required to remove an electron from a gaseous atom. For example, 8.64×10^{-19} J is required to remove an electron from a gaseous lithium atom. The energy required to remove the first outermost electron from an atom is called the first ionization energy. The first ionization energy of lithium equals 8.64×10^{-19} J. The loss of the electron results in the formation of a Li^+ ion. The first ionization energies of the elements in periods 1 through 5 are plotted on the graph in **Figure 16**.

READING CHECK **Define** *ionization energy.*

Think of ionization energy as an indication of how strongly an atom's nucleus holds onto its valence electrons. A high ionization energy value indicates the atom has a strong hold on its electrons. Atoms with large ionization energy values are less likely to form positive ions. Likewise, a low ionization energy value indicates an atom loses an outer electron easily. Such atoms are likely to form positive ions. Lithium's low ionization energy, for example, is important for its use in lithium-ion computer backup batteries, where the ability to lose electrons easily makes a battery that can quickly provide a large amount of electrical power.



5	Describe how ions (cations and anions) form to fulfill the octet rule	Text Book	75-76-77
	يصف تكون الأيونات (كاتيون - أنيون) للوصول إلى الترتيب الإلكتروني المستقر	نص الكتاب	

Positive Ion Formation

A positive ion forms when an atom loses one or more valence electrons in order to attain a noble gas configuration. A positively charged ion is called a **cation.** To understand the formation of a positive ion, compare the electron configurations of the noble gas neon (atomic number 10) and the alkali metal sodium (atomic number 11).

Neon atom (Ne)	$1s^2 2s^2 2p^6$
Sodium atom (Na)	$1s^2 2s^2 2p^6 3s^1$

Note that the sodium atom has one 3s valence electron; it differs from the noble gas neon by that single valence electron. When sodium loses this outer valence electron, the resulting electron configuration is identical to that of neon. **Figure 2** shows how a sodium atom loses its valence electron to become a sodium cation.

By losing an electron, the sodium atom acquires the stable outer-electron configuration of neon. It is important to understand that although sodium now has the electron configuration of neon, it is not neon. It is a sodium ion with a single positive charge. The 11 protons that establish the character of sodium still remain within its nucleus.

Metal ions Metals atoms are reactive because they lose valence electrons easily. The group 1 and 2 metals are the most reactive metals on the periodic table. For example, potassium and magnesium, group 1 and 2 elements, respectively, form K^+ and Mg^{2+} ions. Some group 13 atoms also form ions. The ions formed by metal atoms in groups 1, 2, and 13 are summarized in **Table 2**.

Transition metal ions Recall that, in general, transition metals have an outer energy level of ns^2. Going from left to right across a period, atoms of each element fill an inner d sublevel. When forming positive ions, transition metals commonly lose their two valence electrons, forming 2+ ions. However, it is also possible for d electrons to be lost. Thus, transition metals also commonly form ions of 3+ or greater, depending on the number of d electrons in the electron structure. It is difficult to predict the number of electrons that will be lost. For example, iron (Fe) forms both Fe^{2+} and Fe^{3+} ions. A useful rule of thumb for these metals is that they form ions with a 2+ or a 3+ charge.

Pseudo-noble gas configurations Although the formation of an octet is the most stable electron configuration, other electron configurations can also provide some stability. For example, elements in groups 11–14 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as pseudo-noble gas configurations. In **Figure 3**, the zinc atom has the electron configuration of $1s^22s^22p^63s^23p^64s^23d^{10}$. When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^22s^22p^63s^23p^63d^{10}$ results in a pseudo-noble gas configuration.

Negative Ion Formation

Nonmetals, which are located on the right side of the periodic table, easily gain electrons to attain a stable outer electron configuration. Examine **Figure 4**. To attain a noble-gas configuration, chlorine gains one electron, forming an ion with a 1– charge. After gaining the electron, the chloride ion has the electron configuration of an argon atom.

Chlorine atom (Cl)	$1s^22s^22p^63s^23p^5$
Argon atom (Ar)	$1s^22s^22p^63s^23p^6$
Chloride ion (Cl^-)	$1s^22s^22p^63s^23p^6$

An **anion** is a negatively charged ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus, a chlorine atom becomes a chloride anion. What is the name of the nitrogen anion?

Nonmetal ions As shown in **Table 3**, nonmetals gain the number of electrons that, when added to their valence electrons, equals 8. For example, consider phosphorus, with five valence electrons. To form a stable octet, the atom gains three electrons and forms a phosphide ion with a 3– charge. Likewise, oxygen, with six valence electrons, gains two electrons and forms an oxide ion with a 2– charge.

Some nonmetals can lose or gain other numbers of electrons to form an octet. For example, in addition to gaining three electrons, phosphorus can lose five. However, in general, group 15 elements gain three electrons, group 16 elements gain two, and group 17 elements gain one to achieve an octet.

Select the best answer.

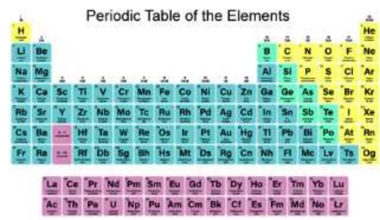
The element magnesium (Mg) is in group 2 on the periodic table.

How do we indicate the ion of this element?

- ○ Mg^{2-}
- ○ Mg^{1+}
- ○ Mg^{1-}
- ◉ Mg^{2+}

Drag and drop the best answer to complete the electron configuration.

Complete the electron configuration of the Fe^{3+} ion.

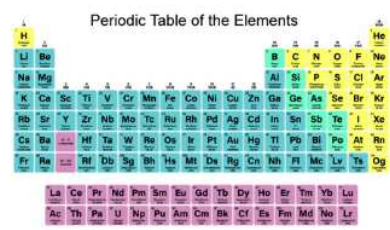


[Ar] $3d^5$ ✓

$4s^2 3d^6 4p^5$ | $4s^2 3d^9$ | $3d^6 4p^7$ | $4s^2 3d^8$ | $4s^2 3d^3$

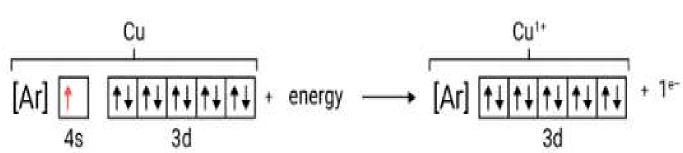
Select the best answer.

Which of these metals would be MOST reactive?



- ○ ^{140}Ce
- ○ ^{27}Al
- ◉ ^{39}K ✓
- ○ ^{48}Ti

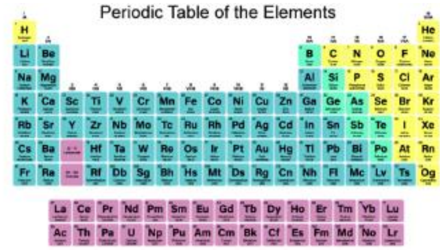
Select the best answer from the drop-down menu to complete the sentence.



When copper loses an electron, the ion has a pseudo-noble gas ✓ configuration.

Select the best answer.

Which elements tend to form ions with a 2+ charge?



- ☑ copper (Cu) ✓
- ☐ chlorine (Cl)
- ☐ sulfur (S)
- ☐ aluminum (Al)
- ☐ potassium (K)
- ☑ magnesium (Mg) ✓

Select the best answer.

Mercury (Hg) is a Group 12 metal with an electron configuration of $[Xe]6s^2\ 4f^{14}\ 5d^{10}$. What is the electron configuration of the 2+ ion of mercury?

- ◉ $[Xe]4f^{14}5d^{10}$ ✓
- ○ $[Xe]4f^{14}5d^{8}$
- ○ $[Xe]6s^{2}4f^{14}5d^{8}$
- ○ $[Xe]6s^{2}4f^{14}$

Select the correct answer from the drop-down menu.

Negatively charged ions are called [anions ✓].

Select the best answer.

Sulfur is an element in group 16. How would we represent the expected ion of sulfur?

- ◉ S^{2-} ✓
- ○ S^{2+}
- ○ S^{1+}
- ○ S^{1-}

Select the best answer.

Which of the following is true about a chloride ion?

- ○ A chloride ion is formed from a calcium atom.
- ○ A chloride ion is positively charged.
- ◉ A chloride ion is negatively charged. ✓
- ○ A chloride ion is a metallic ion.

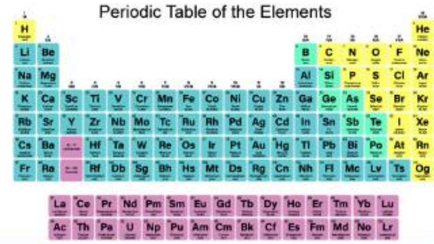
Select the best answer.

The element sulfur (S) has an atomic number of 16 and is found in group 16. What is the electron configuration of a sulfide ion?

- [x] $1s^2 2s^2 2p^6 3s^2 3p^6$ ✓
- [ ] $1s^2 2s^2 2p^6 3s^2 3p^4$
- [ ] $1s^2 2s^2 2p^6$
- [ ] $1s^2 2s^2 2p^6 3s^2$

Select the best answer.

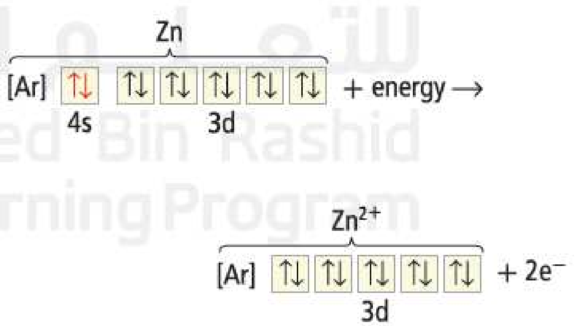
Which of the following elements tend to form ions with a 1- charge?



- [ ] cesium (Cs)
- [x] iodine (I) ✓
- [ ] neon (Ne)
- [ ] calcium (Ca)
- [ ] phosphorus (P)
- [x] chlorine (Cl) ✓

6	Describe how ions (cations and anions) form to fulfill the octet rule يصف تكون الأيونات (كاتيون - أنيون) للوصول إلى الترتيب الإلكتروني المستقر	Text Book نص الكتاب	76

Pseudo-noble gas configurations Although the formation of an octet is the most stable electron configuration, other electron configurations can also provide some stability. For example, elements in groups 11–14 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as pseudo-noble gas configurations. In **Figure 3**, the zinc atom has the electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$. When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ results in a pseudo-noble gas configuration.



When the two 4s valence electrons are lost, a stable pseudo-noble gas configuration consisting of filled s, p, and d sublevels is achieved. Note that the filled 3s and 3p orbitals exist as part of the [Ar] configuration.

how zinc (Zn) and tin (Sn) achieve stability by gaining a pseudo-noble gas configuration.

A zinc atom (Zn) loses 2 valence electrons from its 4s orbital to form a zinc ion (Zn^{2+}).

zinc atom (Zn)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

zinc ion (Zn^{2+})

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

A tin atom (Sn) can lose 4 valence electrons from its 5s and 5p orbitals to form a tin ion (Sn^{4+}).

tin atom (Sn)

$[Kr]\ 5s^2\ 4d^{10}\ 5p^2$

tin ion (Sn^{4+})

$[Kr]\ 4d^{10}$

7	Use the Lewis diagram (electron- dot structure) to explain how elements from the periodic table groups combine to form an ionic compound	Table 4 , Text Book	79-80
	يستخدم الترميز النقطي للإلكترون لتفسير ارتباط عناصر المجموعات المختلفة في الجدول الدوري لتكوين مركب أيوني	الجدول 4 ونص الكتاب	

Table 4 Formation of Sodium Chloride

Chemical Equation

$Na + Cl \longrightarrow Na^+ + Cl^- + \text{energy}$

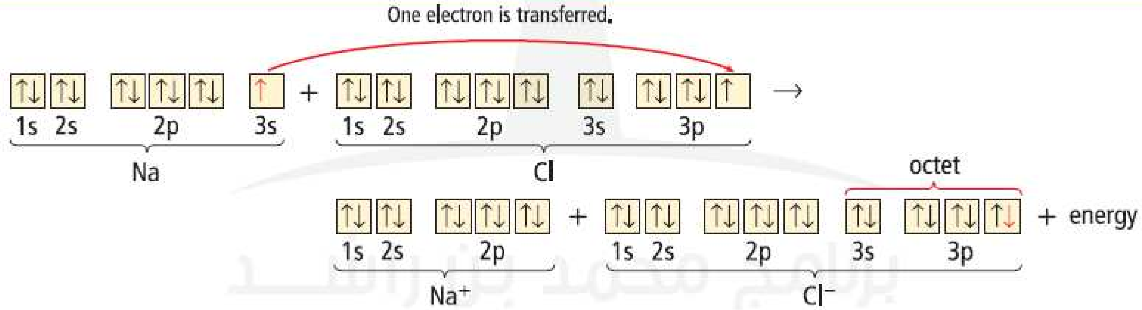
Electron Configurations

One electron is transferred.

$\underbrace{[Ne]3s^1}_{Na} + \underbrace{[Ne]3s^23p^5}_{Cl} \longrightarrow \underbrace{[Ne]}_{Na^+} + \underbrace{[Ar]}_{Cl^-} + \text{energy}$

Orbital Notation

One electron is transferred.

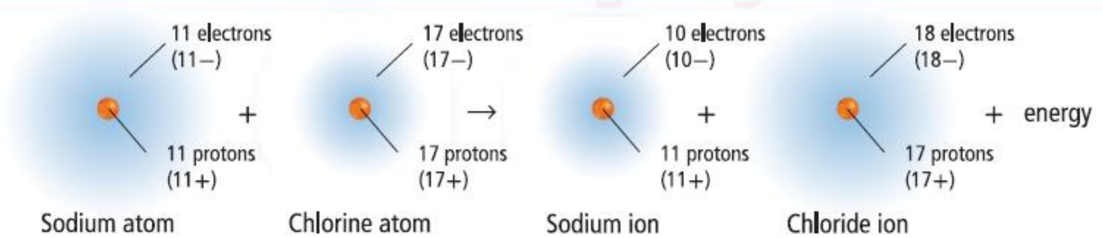


Electron-Dot Structures

One electron is transferred.

$Na\cdot + \cdot\ddot{Cl}: \longrightarrow [Na]^+ + [:\ddot{Cl}:]^- + \text{energy}$

Atomic Models



Compound formation and charge What role does ionic charge play in the formation of ionic compounds? To answer this question, examine how calcium fluoride forms. Calcium has the electron configuration [Ar]$4s^2$, and needs to lose two electrons to attain the stable configuration of argon. Fluorine has the configuration [He]$2s^2 2p^5$, and must gain one electron to attain the stable configuration of neon. Because the number of electrons lost and gained must be equal, two fluorine atoms are needed to accept the two electrons lost from the calcium atom.

$$1\ \text{Ca}\ \cancel{\text{ion}} \left(\frac{2+}{\text{Ca}\ \cancel{\text{ion}}}\right) + 2\ \text{F}\ \cancel{\text{ions}} \left(\frac{1-}{\text{F}\ \cancel{\text{ion}}}\right) = (1)(2+) + (2)(1-) = 0$$

As you can see, the overall charge of one unit of calcium fluoride (CaF_2) is zero. **Table 4** summarizes several ways in which the formation of an ionic compound such as sodium chloride can be represented.

Next, consider aluminum oxide, the whitish coating that forms on aluminum chairs. To acquire a noble-gas configuration, each aluminum atom loses three electrons and each oxygen atom gains two electrons. Thus, three oxygen atoms are needed to accept the six electrons lost by two aluminum atoms. The neutral compound formed is aluminum oxide (Al_2O_3).

$$2\ \text{Al}\ \cancel{\text{ions}} \left(\frac{3+}{\text{Al}\ \cancel{\text{ion}}}\right) + 3\ \text{O}\ \cancel{\text{ions}} \left(\frac{2-}{\text{O}\ \cancel{\text{ion}}}\right) = 2(3+) + 3(2-) = 0$$

8	Describe the relationship between lattice energy and the charge of ions	Table 6 , Text Book	84-85
	يصف العلاقة بين طاقة الشبكة وشحنة الأيون	الجدول 6 ونص الكتاب	

Energy and the Ionic Bond

During every chemical reaction, energy is either absorbed or released. If energy is absorbed during a chemical reaction, the reaction is endothermic. If energy is released, it is exothermic.

The formation of ionic compounds from positive ions and negative ions is always exothermic. The attraction of the positive ion for the negative ions close to it forms a more stable system that is lower in energy than the individual ions. If the amount of energy released during bond formation is reabsorbed, the bonds holding the positive ions and negative ions together will break apart.

Lattice energy Because the ions in an ionic compound are arranged in a crystal lattice, the energy required to separate 1 mol of the ions of an ionic compound is referred to as the **lattice energy.** The strength of the forces holding ions in place is reflected by the lattice energy. The greater the lattice energy, the stronger the force of attraction.

Lattice energy is directly related to the size of the ions bonded. Smaller ions form compounds with more closely spaced ionic charges. Because the electrostatic force of attraction between opposite charges increases as the distance between the charges decreases, smaller ions produce stronger interionic attractions and greater lattice energies. For example, the lattice energy of a lithium compound is greater than that of a potassium compound containing the same anion because the lithium ion is smaller than the potassium ion.

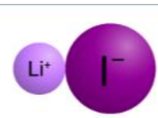
Table 6 Lattice Energies of Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
KI	632	KF	808
KBr	671	AgCl	910
RbF	774	NaF	910
NaI	682	LiF	1030
NaBr	732	$SrCl_2$	2142
NaCl	769	MgO	3795

The value of lattice energy is also affected by the charge of the ion. The ionic bond formed from the attraction of ions with larger positive or negative charges generally has a greater lattice energy. The lattice energy of MgO is almost four times greater than that of NaF because the charge of the ions in MgO is greater than the charge of the ions in NaF. The lattice energy of $SrCl_2$ is between the lattice energies of MgO and NaF because $SrCl_2$ contains ions with both higher and lower charges.

READING CHECK **Summarize** the charges on each ion in the following ionic compounds: MgO, NaF, $SrCl_2$.

Table 6 shows the lattice energies of some ionic compounds. Examine the lattice energies of RbF and KF. Because K^+ has a smaller ionic radius than Rb^+, KF has a greater lattice energy than RbF. This confirms that lattice energy is related to ion size. Notice the lattice energies of $SrCl_2$ and AgCl. How do they show the relationship between lattice energy and the charge of the ions involved?

Compound	Ionic Radii	Lattice Energy (kJ/mol)
lithium fluoride (LiF)		1,036
lithium chloride (LiCl)		853
lithium bromide (LiBr)		807
lithium iodide (LiI)		757

Compound	Charge on Cation	Charge on Anion	Lattice Energy (kJ/mol)
sodium fluoride (NaF)	1+	1–	904
calcium fluoride (CaF_2)	2+	1–	2,611
lithium bromide ($LiBr$)	1+	1–	807
magnesium bromide ($MgBr_2$)	2+	1–	2,421

9	Write the chemical name of an ionic compound containing monoatomic and polyatomic ions(including oxyanions)	Text Book	89-90-91
	يكتب صيغة كيميائية لمركب أيوني يحتوي على أيونات أحادية وأيونات متعددة الذرات(أكسجينية) ويسمي المركبات من خلال صيغتها	نص الكتاب	

Formulas for polyatomic ionic compounds Many ionic compounds contain **polyatomic ions**, which are ions made up of more than one atom. **Table 9** and **Figure 10** list some common polyatomic ions. Also, refer to **Table R-5** in the Student Resources. A polyatomic ion acts as an individual ion in a compound and its charge applies to the entire group of atoms. Thus, the formula for a polyatomic compound follows the same rules used for a binary compound.

Because a polyatomic ion exists as a unit, never change subscripts of the atoms within the ion. If more than one polyatomic ion is needed, place parentheses around the ion and write the appropriate subscript outside the parentheses. For example, consider the compound formed from the ammonium ion (NH_4^+) and the oxide ion (O^{2-}). To balance the charges, the compound must have two ammonium ions for each oxide ion. To add a subscript to ammonium, enclose it in parentheses, then add the subscript. The correct formula is $(NH_4)_2O$.

Names for Ions and Ionic Compounds

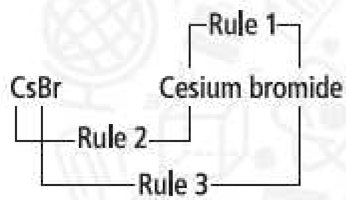
Scientists use a systematic approach when naming ionic compounds. Because ionic compounds have both cations and anions, the naming system accounts for both of these ions.

Naming an oxyanion An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the rules in **Table 10**.

As shown in **Table 11**, chlorine forms four oxyanions that are named according to the number of oxygen atoms present. Names of similar oxyanions formed by other halogens follow the rules used for chlorine. For example, bromine forms the bromate ion (BrO_3^-), and iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Naming ionic compounds Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with chemical formulas, you can use the following five rules to name ionic compounds.

1. Name the cation followed by the anion. Remember that the cation is always written first in the formula.
2. For monatomic cations, use the element name.
3. For monatomic anions, use the root of the element name plus the suffix *-ide*.
   Example:



4. To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation.
   Note: This rule applies to the transition metals and metals on the right side of the periodic table, which often have more than one oxidation number. See **Table 8**. It does not apply to group 1 and group 2 cations, as they have only one oxidation number.
   Examples:
   Fe^{2+} and O^{2-} ions form FeO, known as iron(II) oxide.
   Fe^{3+} and O^{2-} ions form Fe_2O_3, known as iron(III) oxide.
5. When the compound contains a polyatomic ion, simply use the name of the polyatomic ion in place of the anion or cation.
   Examples:
   The name for NaOH is sodium hydroxide.
   The name for $(NH_4)_2S$ is ammonium sulfide.

10	Write the chemical name of an ionic compound containing monoatomic and polyatomic ions(including oxyanions)	Strategy	92
	يكتب صيغة كيميائية لمركب أيوني يحتوي على أيونات أحادية وأيونات متعددة الذرات(أكسجينية) ويسمي المركبات من خلال صيغتها	استراتيجية	

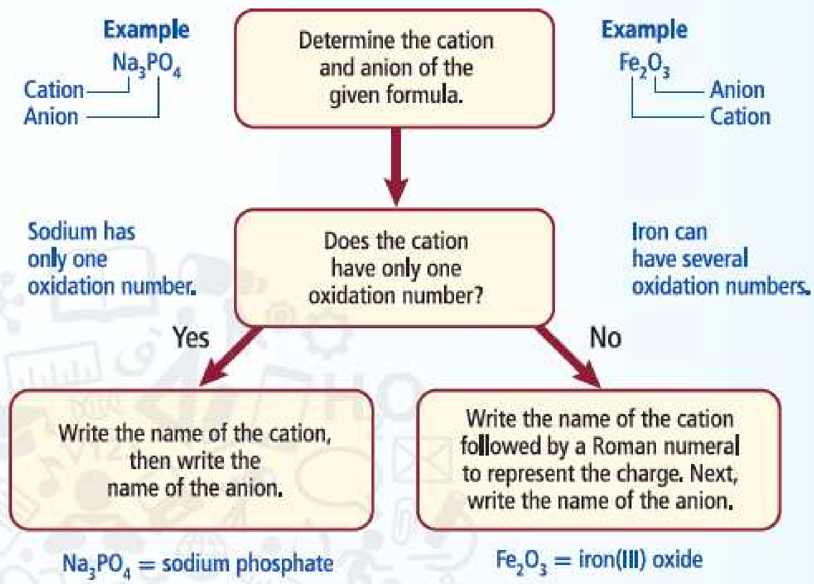
PROBLEM-SOLVING STRATEGY

Naming Ionic Compounds

Naming ionic compounds is easy if you follow this naming-convention flowchart.

Apply the Strategy

Name the compounds KOH and Ag_2CrO_4 using this flowchart.



ملاحظة : يرجى الاطلاع على دروس 48 و 49 و 50 في منصة الف وحل الأمثلة الواردة

11	Explain some physical properties of metals (Melting and boiling points, Thermal and electrical conductivity,Malleability,Ductility,Durability,Hardness,and Strength)	Figures 11 , 12	93 – 94
	يفسر بعض الخواص الفلزية (درجتا الانصهار والغليان– توصيل الحرارة والكهرباء – قابلية الطرق والسحب والمتانة– الصلابة والقوة)	الأشكال 11 و 12	

Figure 11 The valence electrons in metals (shown as a blue cloud of minus signs) are evenly distributed among the metallic cations (shown in red). Attractions between positive cations and the negative "sea" hold the metal atoms together in a lattice.

Explain *Why are electrons in metals known as delocalized electrons?*

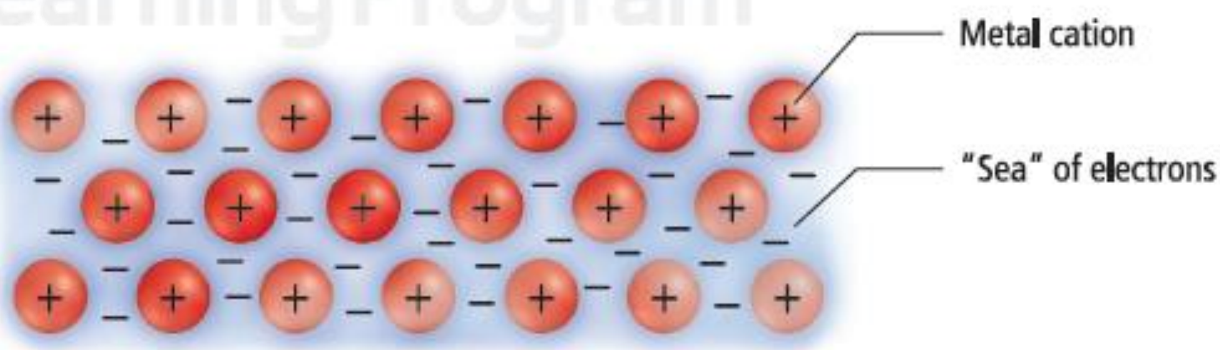
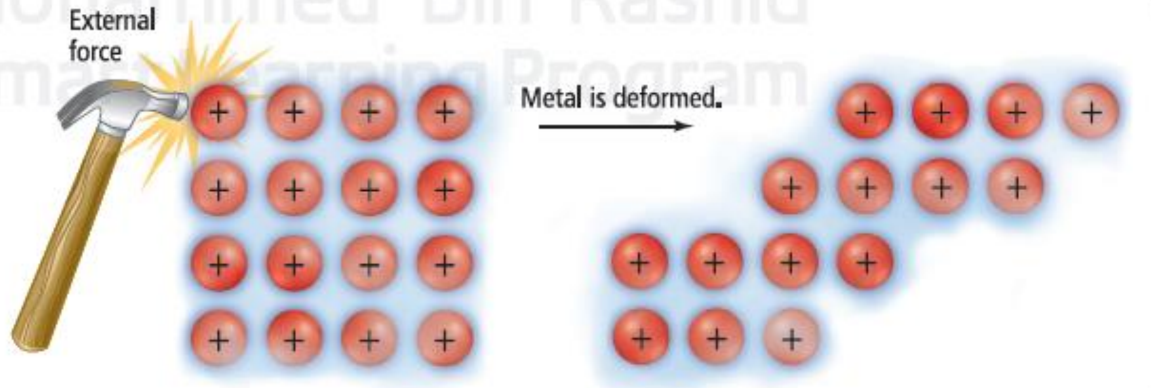


Figure 12 An applied force causes metal ions to move through delocalized electrons, making metals malleable and ductile.



12	Differentiate between sigma and pi bonds	Text Book	112-113-114
	يفرق بين روابط سيجما وروابط باي	نص الكتاب	

The sigma bond Single covalent bonds are also called **sigma bonds,** represented by the Greek letter sigma (σ). A sigma bond occurs when the pair of shared electrons is in an area centered between the two atoms. When two atoms share electrons, their valence atomic orbitals overlap end-to-end, concentrating the electrons in a bonding orbital between the two atoms. A bonding orbital is a localized region where bonding electrons will most likely be found. Sigma bonds can form when an s orbital overlaps with another s orbital or a p orbital, or two p orbitals overlap end-to-end. Water (H_2O), ammonia (NH_3), and methane (CH_4) have sigma bonds, as shown in **Figure 7.**

READING CHECK **List** the orbitals that can form sigma bonds in a covalent compound.

Multiple Covalent Bonds

In some molecules, atoms have noble-gas configurations when they share more than one pair of electrons with one or more atoms. Sharing multiple pairs of electrons forms multiple covalent bonds. A double covalent bond and a triple covalent bond are examples of multiple bonds. Carbon, nitrogen, oxygen, and sulfur atoms often form multiple bonds with other nonmetals. How do you know if two atoms will form a multiple bond? In general, the number of valence electrons needed to form an octet equals the number of covalent bonds that can form.

Double bonds A double covalent bond forms when two pairs of electrons are shared between two atoms. For example, atoms of the element oxygen only exist as diatomic molecules. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration, as shown in **Figure 8a.** A double covalent bond forms when each oxygen atom shares two electrons; a total of two pairs of electrons are shared between the two atoms.

Triple bonds A triple covalent bond forms when three pairs of electrons are shared between two atoms. Diatomic nitrogen (N_2) molecules contain a triple covalent bond. Each nitrogen atom shares three electron pairs, forming a triple bond with the other nitrogen atom as shown in **Figure 8b.**

The pi bond A multiple covalent bond consists of one sigma bond and at least one pi bond. A **pi bond,** represented by the Greek letter pi (π), forms when parallel orbitals overlap and share electrons. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together.

The Strength of Covalent Bonds

Recall that a covalent bond involves attractive and repulsive forces. In a molecule, nuclei and electrons attract each other, but nuclei repel other nuclei, and electrons repel other electrons. When this balance of forces is upset, a covalent bond can be broken. Because covalent bonds differ in strength, some bonds break more easily than others. Several factors influence the strength of covalent bonds.

Bond length The strength of a covalent bond depends on the distance between the bonded nuclei. The distance between the two bonded nuclei at the position of maximum attraction is called bond length, as shown in **Figure 10.** It is determined by the sizes of the two bonding atoms and how many electron pairs they share. Bond lengths for molecules of fluorine (F_2), oxygen (O_2), and nitrogen (N_2) are listed in **Table 1.** Notice that as the number of shared electron pairs increases, the bond length decreases.

Bond length and bond strength are also related: the shorter the bond length, the stronger the bond. Therefore, a single bond, such as that in F_2, is weaker than a double bond, such as that in O_2. Likewise, the double bond in O_2 is weaker than the triple bond in N_2.

13	Identify the relationship between the type of a covalent bond (single, double, triple) and its bond length, bond strength and bond dissociation energy	Tables 1 , 2 , Text Book	114 - 115
	يحدد العلاقة بين طول الرابطة التساهمية (أحادية – ثنائية – ثلاثية) وقوتها وطاقة تفككها	الجداول 1 و 2 ونص الكتاب	

Table 1 Covalent Bond Type and Bond Length

Molecule	Bond Type	Bond Length
F_2	single covalent	1.43×10^{-10} m
O_2	double covalent	1.21×10^{-10} m
N_2	triple covalent	1.10×10^{-10} m

Table 2 **Bond-Dissociation Energy**

Molecule	Bond-Dissociation Energy
F_2	159 kJ/mol
O_2	498 kJ/mol
N_2	945 kJ/mol

Bonds and energy An energy change occurs when a bond between atoms in a molecule forms or breaks. Energy is released when a bond forms, but energy must be added to break a bond. The amount of energy required to break a specific covalent bond is called bond-dissociation energy and is always a positive value. The bond-dissociation energies for the covalent bonds in molecules of fluorine, oxygen, and nitrogen are listed in **Table 2.**

Bond-dissociation energy also indicates the strength of a chemical bond because of the inverse relationship between bond energy and bond length. As indicated in **Table 1** and **Table 2,** the smaller the bond length is, the greater the bond-dissociation energy. The sum of the bond-dissociation energy values for all of the bonds in a molecule is the amount of chemical potential energy in a molecule of that compound.

The total energy change of a chemical reaction is determined from the energy of the bonds broken and formed. An **endothermic reaction** occurs when a greater amount of energy is required to break the existing bonds in the reactants than is released when the new bonds form in the products. An **exothermic reaction** occurs when more energy is released during product bond formation than is required to break bonds in the reactants. **Figure 11** illustrates a common exothermic reaction. You will study exothermic and endothermic reactions in much greater detail when you study the energy changes in chemical reactions.

14	Name a binary molecular compound based on its molecular formula	Example 2 , Applications	117
	أيسمي مركب جزيئي ثنائي من صيغته الجزيئية	مثال 2 وتطبيقات	

EXAMPLE 2

NAMING BINARY MOLECULAR COMPOUNDS Name the compound P_2O_5, which is used as a drying and dehydrating agent.

1 ANALYZE THE PROBLEM

You are given the formula for a compound. The formula contains the elements and the number of atoms of each element in one molecule of the compound. Because only two different elements are present and both are nonmetals, the compound can be named using the rules for naming binary molecular compounds.

2 SOLVE FOR THE UNKNOWN

First, name the elements involved in the compound.

phosphorus — **The first element, represented by P, is phosphorus.**

oxide — **The second element, represented by O, is oxygen. Add the suffix *–ide* to the root of oxygen, *ox-*.**

phosphorus oxide — **Combine the names.**

Now modify the names to indicate the number of atoms present in a molecule.

diphosphorus pentoxide — **From the formula P_2O_5, you know that two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From Table 3, you know that *di-* is the prefix for two and *penta-* is the prefix for five. The *a* in *penta-* is not used because *oxide* begins with a vowel.**

Name each of the binary covalent compounds listed below.

14. CO_2
15. SO_2
16. NF_3
17. CCl_4
18. **Challenge** What is the formula for diarsenic trioxide?

Table 3 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-

15	يُسمي المحاليل الحمضية(الثنائية - الأكسجينية) بإعطاء صيغتها الكيميائية والعكس	Tables4 , Text Book	118
	Name an acid (binary acid and oxyacid) given its chemical formula and vice versa	الجدول 4 وبنص الكتاب	

Naming Acids

Water solutions of some molecules are acidic and are named as acids. Acids are important compounds with specific properties. If a compound produces hydrogen ions (H^+) in solution, it is an acid. For example, HCl produces H^+ in solution and is an acid. Two common types of acids exist–binary acids and oxyacids.

Naming binary acids A binary acid contains hydrogen and one other element. The naming of the common binary acid known as hydrochloric acid is explained in the following rules.

1. The first word has the prefix *hydro-* to name the hydrogen part of the compound. The rest of the first word consists of a form of the root of the second element plus the suffix *-ic*. HCl (hydrogen and chlorine) becomes *hydrochloric*.
2. The second word is always *acid*. Thus, HCl in a water solution is called *hydrochloric acid*.

Although the term *binary* indicates exactly two elements, a few acids that contain more than two elements are named according to the rules for naming binary acids. If no oxygen is present in the formula for the acidic compound, the acid is named in the same way as a binary acid, except that the root of the second part of the name is the root of the polyatomic ion that the acid contains. For example, HCN, which is composed of hydrogen and the cyanide ion, is called *hydrocyanic acid* in solution.

Naming oxyacids An acid that contains both a hydrogen atom and an oxyanion is referred to as an **oxyacid.** Recall that an oxyanion is a polyatomic ion containing one or more oxygen atoms. The following rules explain the naming of nitric acid (HNO_3), an oxyacid.

1. First, identify the oxyanion present. The first word of an oxyacid's name consists of the root of the oxyanion and the prefix *per-* or *hypo-* if it is part of the oxyanion's name. The first word of the oxyacid's name also has a suffix that depends on the oxyanion's suffix. If the oxyanion's name ends with the suffix *-ate*, replace it with the suffix *-ic*. If the name of the oxyanion ends with the suffix *-ite*, replace it with the suffix *-ous*. NO_3^-, the nitrate ion, becomes *nitric*.
2. The second word of the name is always *acid*. HNO_3 (hydrogen and the nitrate ion) becomes *nitric acid*.

Table 4 Naming Oxyacids

Compound	Oxyanion	Acid Suffix	Acid Name
$HClO_3$	chlorate	-ic	chloric acid
$HClO_2$	chlorite	-ous	chlorous acid
HNO_3	nitrate	-ic	nitric acid
HNO_2	nitrite	-ous	nitrous acid

16	Draw Lewis structures for a number of covalent compounds with single and multiple bonds	Example 3 , Applications	123
	يرسم بنى لويس لعدد من المركبات التساهمية ذات الروابط الأحادية والمتعددة	مثال 3 وتطبيقات	

EXAMPLE 3

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH SINGLE BONDS Ammonia is a raw material used in the manufacture of many products, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

1 ANALYZE THE PROBLEM

Ammonia molecules consist of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}}$$
$$= 8 \text{ valence electrons}$$

There are 8 valence electrons available for bonding.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs. To do this, divide the number of available electrons by two.

Four pairs of electrons are available for bonding.

```
H—N—H
  |
  H
```

Place a bonding pair (a single bond) between the central nitrogen atom and each terminal hydrogen atom.

Determine the number of bonding pairs remaining.

4 pairs total − 3 pairs used
= 1 pair available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

The remaining pair—a lone pair—must be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs.

```
   ..
H—N—H
  |
  H
```

Place the remaining lone pair on the central nitrogen atom.

APPLICATIONS

37. Draw the Lewis structure for BH_3.

38. **Challenge** A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

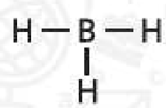
17	Determine the exceptions to the octet rule (odd number of valence electrons, suboctets and coordinate covalent bonds, expanded octets)	Text Book	126-127-128
	أحدد الجزيئات التي تُعتبر استثناءات لقاعدة الثمانية (العدد الفردي لإلكترونات التكافؤ – الثمانيات الفرعية والروابط التساهمية التناسقية والثمانيات الموسعة)	نص الكتاب	

Exceptions to the Octet Rule

Generally, atoms attain an octet when they bond with other atoms. Some molecules and ions, however, do not obey the octet rule. There are several reasons for these exceptions.

Odd number of valence electrons First, a small group of molecules might have an odd number of valence electrons and be unable to form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs. See **Figure 15.** ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.

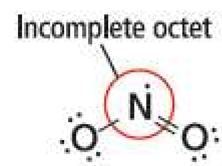
Subsoctets and coordinate covalent bonds Another exception to the octet rule is due to a few compounds that form suboctets–stable configurations with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 13 metalloid, forms three covalent bonds with other nonmetallic atoms.



The boron atom shares only six electrons–too few to form an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom.

A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons to form a stable electron arrangement with lower potential energy. Refer to **Figure 16.** Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

Figure 15 The central nitrogen atom in this NO_2 molecule does not satisfy the octet rule; the nitrogen atom has only seven electrons in its outer energy level.

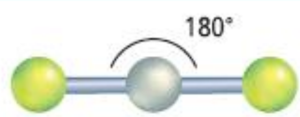
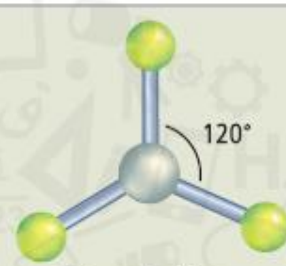
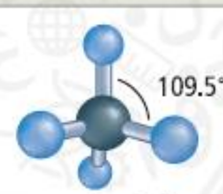


Expanded octets The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an expanded octet. An expanded octet can be explained by considering the d orbitals that occur in the energy levels of elements in period three or higher. An example of an expanded octet, shown in **Figure 17**, is the bond formation in the molecule PCl_5. Five bonds are formed with ten electrons shared in one s orbital, three p orbitals, and one d orbital. Another example is the molecule SF_6, which has six bonds sharing 12 electrons in an s orbital, three p orbitals, and two d orbitals. When you draw the Lewis structures for these compounds, either extra lone pairs are added to the central atom or more than four bonding atoms are present in the molecule.

ملاحظة : يرجى الاطلاع على درس 68 في منصة الف وحل الأمثلة الواردة لكل استثناء لقاعدة الثمانية

18	Using the VSEPR model the electron domain geometry and molecular geometry for different molecules and ions	Tables 6 , Text Book	131
	استعمال نموذج VSEPR لتوقع شكل وزوايا الربط في الجزيئات والأيونات المختلفة	الجدول 6 ونص الكتاب	

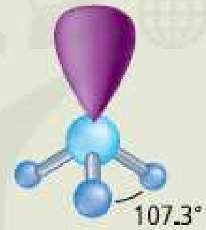
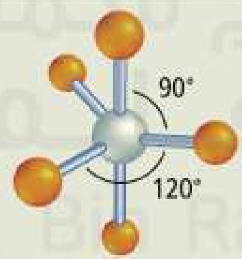
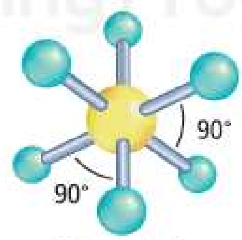
Table 6 Molecular Shapes

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
$BeCl_2$	2	2	0	sp	 Linear
$AlCl_3$	3	3	0	sp^2	 Trigonal planar
CH_4	4	4	0	sp^3	 Tetrahedral

The $BeCl_2$ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in $AlCl_3$ have maximum separation in a trigonal planar shape with 120° bond angles.

When the central atom in a molecule has four pairs of bonding electrons, as CH_4 does, the shape is tetrahedral. The bond angles are 109.5°.

NH_3	4	3	1	sp^3	 **Trigonal pyramidal**	NH_3 has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.
H_2O	4	2	2	sp^3	 **Bent**	Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramid. As a result, water molecules have a bent shape.
$NbBr_5$	5	5	0	sp^3d	 **Trigonal bipyramidal**	The $NbBr_5$ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.
SF_6	6	6	0	sp^3d^2	 **Octahedral**	As with $NbBr_5$, SF_6 has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

ملاحظة : يرجى الاطلاع على درس 69 في منصة الف وحل الأمثلة الواردة

19	Categorize bond types using electronegativity difference into mostly ionic, polar covalent, mostly covalent, and non– polar covalent bonds	Tables 7, Figures 21, 22, Text Book	134 - 135
	يُصنف الروابط بناءً على الفرق في السالبية الكهربائية إلى (أيونية غالبًا – تساهمية قطبية – تساهمية غالبًا – تساهمية غير قطبية)	الجدول 7 والأشكال 21 و 22 ونص الكتاب	

Table 7 EN Difference and Bond Character

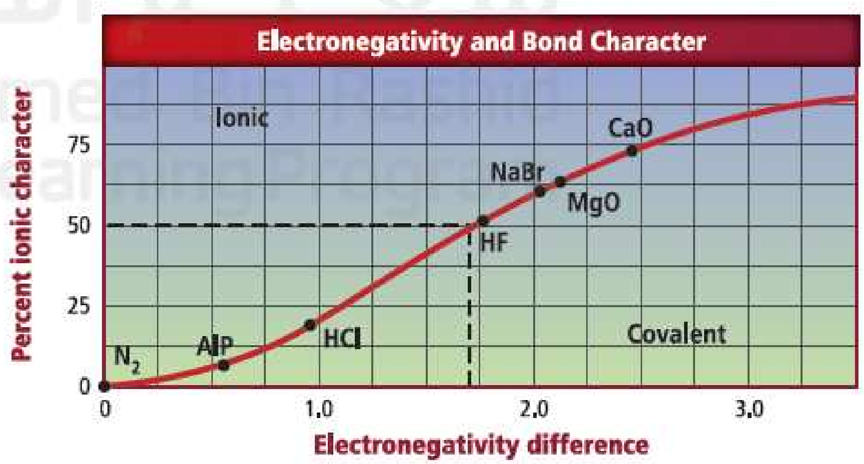
Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 — 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

Electronegativity Values for Selected Elements

Metal / Metalloid / Nonmetal

1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1														

■ **Figure 21** This graph shows that the difference in electronegativity between bonding atoms determines the percent ionic character of the bond. Above 50% ionic character, bonds are mostly ionic.



GRAPH CHECK

Determine the percent ionic character of calcium oxide.

Bond character A chemical bond between atoms of different elements is never completely ionic or covalent. The character of a bond depends on how strongly each of the bonded atoms attracts electrons. As shown in **Table 7**, the character and type of a chemical bond can be predicted using the electronegativity difference of the elements that bond. Electrons in bonds between identical atoms have an electronegativity difference of zero—meaning that the electrons are equally shared between the two atoms. This type of bond is considered nonpolar covalent, or a pure covalent bond. On the other hand, because different elements have different electronegativities, the electron pairs in a covalent bond between different atoms are not shared equally. Unequal sharing results in a **polar covalent bond.** When there is a large difference in the electronegativity between bonded atoms, an electron is transferred from one atom to the other, which results in bonding that is primarily ionic.

Bonding is not often clearly ionic or covalent. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. As the difference in electronegativity increases, the bond becomes more ionic in character. Generally, ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together. **Figure 21** summarizes the range of chemical bonding between two atoms. What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? Where would LiBr be plotted on the graph?

Polar Covalent Bonds

As you just learned, polar covalent bonds form because not all atoms that share electrons attract them equally. A polar covalent bond is similar to a tug-of-war in which the two teams are not of equal strength. Although both sides share the rope, the stronger team pulls more of the rope toward its side. When a polar bond forms, the shared electron pair or pairs are pulled toward one of the atoms. Thus, the electrons spend more time around that atom than the other atom. This results in partial charges at the ends of the bond.

The Greek letter delta (δ) is used to represent a partial charge. In a polar covalent bond, δ^- represents a partial negative charge and δ^+ represents a partial positive charge. As shown in **Figure 22**, δ^- and δ^+ can be added to a molecular model to indicate the polarity of the covalent bond. The more-electronegative atom is at the partially negative end, while the less-electronegative atom is at the partially positive end. The resulting polar bond often is referred to as a dipole (two poles).

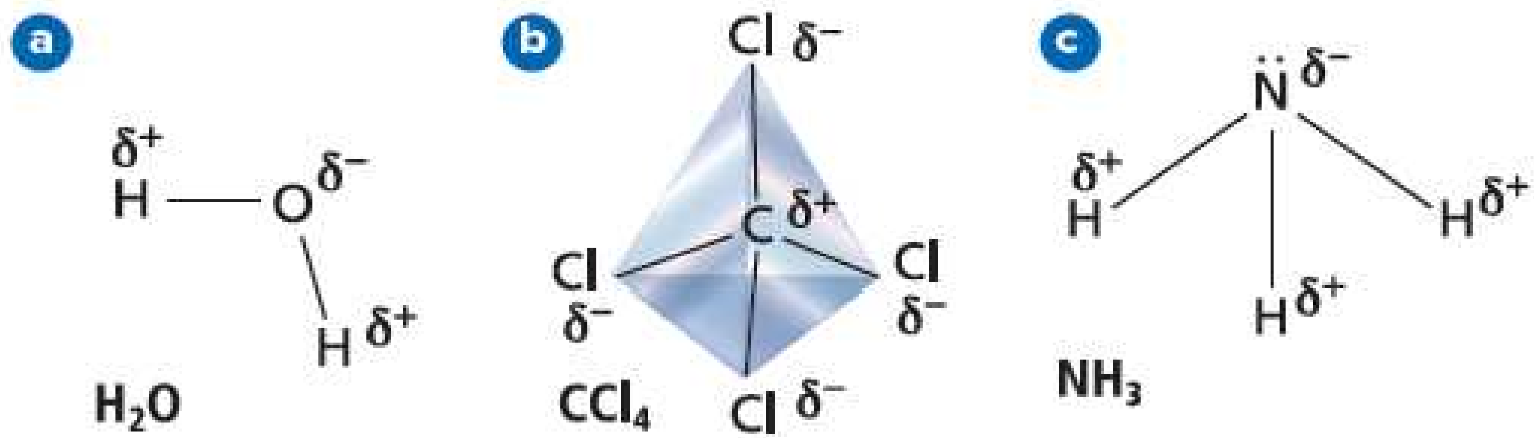
Molecular polarity Covalently bonded molecules are either polar or nonpolar; which type depends on the location and nature of the covalent bonds in the molecule. A distinguishing feature of nonpolar molecules is that they are not attracted by an electric field. Polar molecules, however, are attracted by an electric field. Because polar molecules are dipoles with partially charged ends, they have an uneven electron density. This results in the tendency of polar molecules to align with an electric field.

Polarity and molecular shape You can learn why some molecules are polar and some are not by comparing water (H_2O) and carbon tetrachloride (CCl_4) molecules. Both molecules have polar covalent bonds. According to the data in **Figure 20,** the electronegativity difference between a hydrogen atom and an oxygen atom is 1.24. The electronegativity difference between a chlorine atom and a carbon atom is 0.61. Although these electronegativity differences vary, a H–O bond and a C–Cl bond are considered to be polar covalent.

$$\overset{\delta^+}{\mathrm{H}}\text{–}\overset{\delta^-}{\mathrm{O}} \qquad \overset{\delta^+}{\mathrm{C}}\text{–}\overset{\delta^-}{\mathrm{Cl}}$$

According to their molecular formulas, both molecules have more than one polar covalent bond. However, only the water molecule is polar. Why might one molecule with polar covalent bonds be polar, while a second molecule with polar covalent bonds is nonpolar? The answer lies in the shapes of the molecules.

20	Describe conditions needed for a molecule to be polar	Figure 23 ,Text Book	136
	حدد الشروط التي يجب توافرها ليكون المركب الجزيئي قطبيًا	الشكل 23 ونص الكتاب	



The bent shape of a water molecule makes it polar.

The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

The asymmetric shape of an ammonia molecule results in an unequal charge distribution and the molecule is polar.

The shape of an H_2O molecule, as determined by VSEPR, is bent because the central oxygen atom has lone pairs of electrons, as shown in **Figure 23a.** Because the polar H–O bonds are asymmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.

A CCl_4 molecule is tetrahedral, and therefore, symmetrical, as shown in **Figure 23b.** The electric charge measured at any distance from its center is identical to the charge measured at the same distance to the opposite side. The average center of the negative charge is located on each chlorine atom. The positive center is also located on the carbon atom. Because the partial charges are balanced, CCl_4 is a nonpolar molecule. Note that symmetric molecules are usually nonpolar, and molecules that are asymmetric are polar as long as the bond type is polar.

Is the molecule of ammonia (NH_3), shown in **Figure 23c,** polar? It has a central nitrogen atom and three terminal hydrogen atoms. Its shape is trigonal pyramidal because of the lone pair of electrons present on the nitrogen atom. Using **Figure 20,** you can find that the electronegativity difference of hydrogen and nitrogen is 0.84, making each N–H bond polar covalent. The charge distribution is unequal because the molecule is asymmetric. Thus, the molecule is polar.

Solubility of polar molecules The physical property known as solubility is the ability of a substance to dissolve in another substance. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds are usually soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances, as shown in **Figure 24.**

Figure 24 Symmetric covalent molecules, such as oil and most petroleum products, are nonpolar. Asymmetric molecules, such as water, are usually polar. As shown in this photo, polar and nonpolar substances usually do not mix.

Infer ***Will water alone clean oil from a fabric?***

21	A learning outcome from the SoW****	Undisclosed	Undisclosed
	****ناتج من الخطة الفصلية	غير معلن	غير معلن
22	A learning outcome from the SoW	Undisclosed	Undisclosed
	ناتج من الخطة الفصلية	غير معلن	غير معلن
23	A learning outcome from the SoW	Undisclosed	Undisclosed
	ناتج من الخطة الفصلية	غير معلن	غير معلن
24	A learning outcome from the SoW	Undisclosed	Undisclosed
	ناتج من الخطة الفصلية	غير معلن	غير معلن
25	A learning outcome from the SoW	Undisclosed	Undisclosed
	ناتج من الخطة الفصلية	غير معلن	غير معلن

ملاحظة : نواتج التعلم من 21 الى 25 هي عبارة عن أسئلة إضافية

مع تمنياتي لكم بالتوفيق والنجاح

الأستاذ عدي العاصي