Section 7.1 Ion Formation
pages 206–209

Section 7.1 Assessment
page 209
1. Compare the stability of a lithium atom with that of its ion, Li⁺.
   The Li⁺ ion is more stable because it has a complete octet.

2. Describe two different causes of the force of attraction in a chemical bond.
   The two causes are the attraction between the positive nucleus of one atom and the negative electrons of another atom, and the attraction between positive ions and negative ions.

3. Apply Why are all of the elements from group 18 relatively unreactive, whereas those in group 17 are very reactive?
   The group 18 elements, known as noble gases, have complete outer energy levels and do not easily form ions. The group 17 elements are highly reactive because they need only to gain a single electron to form an octet.

4. Summarize ionic bond formation by correctly pairing these terms: cation, anion, electron gain, electron loss.
   (anion, electron gain), (cation, electron loss)

5. Apply Write out the electron configuration for each atom. Then predict the change that must occur in each to achieve a noble gas configuration.
   a. nitrogen
      \[
      [\text{He}]2s^22p^3; \text{gain 3 electrons (3}^- \text{ion) or lose 5 electrons (5}^+ \text{ion)}
      \]
   b. sulfur
      \[
      [\text{Ne}]3s^23p^4; \text{gain 2 electrons (2}^- \text{ion)}
      \]
   c. barium
      \[
      [\text{Xe}]6s^2; \text{lose 2 electrons (2}^+ \text{ion)}
      \]
   d. lithium
      \[
      [\text{He}]2s^1; \text{lose 1 electron (1}^+ \text{ion)}
      \]

6. Model Draw models to represent the formation of the positive calcium ion and the negative bromide ion.
   Models should show that the calcium atom loses two electrons, forming Ca²⁺, and that bromine gains one electron, forming Br⁻. The models should also show the addition of energy to form the calcium ion, and the release of energy to form the bromide ion.

Section 7.2 Ionic Bonds and Ionic Compounds
pages 210–217

Practice Problems
page 212
Explain how an ionic compound forms from these elements.
7. sodium and nitrogen
   Three Na atoms each lose 1 e⁻, forming 1⁺ ions.
   One N atom gains 3 e⁻, forming a 3⁻ ion. The ions attract, forming Na₃N.
   \[
   3 \text{Na}^{+} + 1 \text{N}^{3-} \rightarrow \text{Na}_3\text{N}
   \]
   The overall charge on one formula unit of Na₃N is zero.

8. lithium and oxygen
   Two Li atoms each lose 1 e⁻, forming 1⁺ ions.
   One O atom gains 2 e⁻, forming a 2⁻ ion. The ions attract, forming Li₂O.
   \[
   2 \text{Li}^{+} + 1 \text{O}^{2-} \rightarrow \text{Li}_2\text{O}
   \]
   The overall charge on one formula unit of Li₂O is zero.
9. strontium and fluorine

One Sr atom loses 2 e\(^{-}\), forming a 2\(^{+}\) ion. Two F atoms each gain 1 e\(^{-}\), forming 1\(^{-}\) ions. The ions attract, forming SrF\(_2\).

\[
\text{1 Sr ion (2}^{+}\text{ Sr ion) + 2 F ions (1}^{-}\text{ F ion)}
\]
\[
= 1(2^{+}) + 2(1^{-}) = 0
\]

The overall charge on one formula unit of SrF\(_2\) is zero.

10. aluminum and sulfur

Two Al atoms each lose 3 e\(^{-}\), forming 3\(^{+}\) ions.

Three S atoms gain 2 e\(^{-}\) each, forming 2\(^{-}\) ions.

The ions attract, forming Al\(_2\)S\(_3\).

\[
\text{2 Al ions (3}^{+}\text{ Al ions) + 3 S ions (2}^{-}\text{ S ion)}
\]
\[
= 2(3^{+}) + 3(2^{-}) = 0
\]

The overall charge on one formula unit of Al\(_2\)S\(_3\) is zero.

11. Challenge

Explain how elements in the two groups shown on the periodic table at the right combine to form an ionic compound.

Three group 1 atoms lose 1 e\(^{-}\), forming 1\(^{+}\) ions.

One group 15 atom gains 3 e\(^{-}\), forming a 3\(^{-}\) ion.

The ions attract, forming X\(_3\)Y, where X represents a group 1 atom and Y represents a group 15 atom.

Section 7.2 Assessment

page 217

12. Explain how an ionic compound made up of charged particles can be electrically neutral.

The total positive charge of the cations in the compound equals the total negative charge of the anions in the compound.

13. Describe the energy change associated with ionic bond formation, and relate it to stability.

Ionic bond formation is exothermic; the lower-energy product is more stable than the original reactants.

14. Identify three physical properties of ionic compounds that are associated with ionic bonds, and relate them to bond strength.

Ionic compounds exist as crystals, have high melting and boiling points, and are hard, rigid, and brittle. They are conductive when dissolved or molten but not when solid. All of these properties are due to the strength of the electrostatic attraction of oppositely-charged ions in close proximity.

15. Explain how ions form bonds and describe the structure of the resulting compound.

Electrons are transferred between atoms forming ions. Electrostatic forces hold the ions together in the ionic compound. The ions are arranged in a regular repeating pattern in an ionic crystal.

16. Relate lattice energy to ionic-bond strength.

As lattice energy becomes more negative, the stronger is the attraction between the ions and, thus, the stronger the ionic bond.

17. Apply Use electron configurations, orbital notation, and electron-dot structures to represent the formation of an ionic compound from the metal strontium and the nonmetal chlorine.

Drawing should include one Sr atom losing 2 e\(^{-}\) and forming an Sr\(^{2+}\) ion, and two Cl atoms each gaining 1 e\(^{-}\) and forming two Cl\(^{-}\) ions. These ions attract, forming SrCl\(_2\).
18. **Design** a concept map that shows the relationships among ionic bond strength, physical properties of ionic compounds, lattice energy, and stability.

Concept maps will vary but should correlate greater bond strength to increased stability and a more negative lattice energy, and that physical properties such as high melting and boiling points, brittleness, and conductivity are due to the strength of ionic bonds.

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**Section 7.3 Names and Formulas for Ionic Compounds**

**Practice Problems**

Write formulas for the ionic compounds formed by the following ions.

19. Potassium and iodide

   KI

20. Magnesium and chloride

   MgCl₂

21. Aluminum and bromide

   AlBr₃

22. Cesium and nitride

   Cs₃N

23. **Challenge** Write the general formula for the ionic compound formed by elements from the two groups shown on the periodic table at the right.

The general formula is XY₂, where X represents the group 2 element and Y represents the group 17 element.

Write formulas for ionic compounds composed of the following ions.

24. Sodium and nitrate

   NaNO₃
25. calcium and chlorate
Ca(ClO₃)₂

26. aluminum and carbonate
Al₂(CO₃)₃

27. Challenge Write the formula for an ionic compound formed by ions from a group 2 element and polyatomic ions composed of only carbon and oxygen.
The polyatomic ion is carbonate (CO₃²⁻).
The general formula is XCO₃, where X is the symbol of a group 2 element.

28. NaBr
sodium bromide

29. CaCl₂
calcium chloride

30. KOH
potassium hydroxide

31. Cu(NO₃)₂
copper(II) nitrate

32. Ag₂CrO₄
silver chromate

33. Challenge The ionic compound NH₄ClO₄ is key reactant used in solid rocket boosters, such as those that power the space shuttle into orbit. Name this compound.
ammonium perchlorate

Section 7.3 Assessment
page 224

34. State the order in which the ions associated with a compound composed of potassium and bromine would be written in the chemical formula and the compound name.
The cation (potassium) is stated first, followed by the anion (bromide).

35. Describe the difference between a monatomic ion and a polyatomic ion and give an example of each.
Monatomic ions are one-atom ions (example, Cl⁻); polyatomic ions are two or more atoms grouped together having a net charge (example, ClO₃⁻).

36. Apply Ion X has a charge of 2⁺ and ion Y has a charge of 1⁻. Write the formula unit of the compound formed from the ions.
XY₂

37. State the name and formula for the compound formed from Mg and Cl.
magnesium chloride, MgCl₂

38. Write the name and formula for the compound formed from sodium ions and nitrite ions.
sodium nitrite, NaNO₂

39. Analyze What subscripts would you most likely use if the following substances formed an ionic compound?
   a. an alkali metal and a halogen
      1, 1
   b. an alkali metal and a nonmetal from group 16
      2, 1
   c. an alkaline earth metal and a halogen
      1, 2
   d. an alkaline earth metal and a nonmetal from group 16
      1, 1
Section 7.4 Metallic Bonds and the Properties of Metals

pages 225–228

Section 7.4 Assessment

pages 228

40. Contrast the structures of ionic compounds and metals.

Ions in ionic compounds are arranged in a repeating pattern of alternating charges, whereas metals consist of fixed cations surrounded by a sea of mobile, or delocalized, electrons.

41. Explain how the conductivity of electricity and the high melting points of metals are explained by metallic bonding.

Delocalized electrons can move through the solid to conduct an electric current. The number of delocalized electrons and the strength of a metallic bond determine the melting point.

42. Contrast the cause of the attraction in ionic bonds and metallic bonds.

Ionic bonds are held together by the electrostatic force of attraction between ions, whereas a metallic bond is due to the attraction of metallic cations for delocalized electrons.

43. Summarize alloy types by correctly pairing these terms and phrases: substitutional, interstitial, replaced, filled in.

(substitutional, replaced), (interstitial, filled in)

44. Design an Experiment Describe an experiment that could be used to distinguish between a metallic solid and an ionic solid. Include at least two different methods for comparing the solids. Explain your reasoning.

A typical student experiment might have steps similar to the following:

1. Test each solid with a conductivity tester.
2. Place the solid in water to determine if it forms a solution.
3. Test the solution with a conductivity tester to determine if it conducts an electric current.
4. Use a hammer to strike the solid and record your observations.

Metallic solids conduct an electric current in the solid state, whereas ionic compounds do not. Metals may react with water, but they do not dissolve in water. Solutions containing ionic compounds conduct an electric current. Metallic solids are malleable and will deform when struck with a hammer, whereas ionic compounds are brittle and will break into pieces when struck with a hammer.

45. Model Draw a model to represent the physical property of metals known as ductility, or the ability to be drawn into a wire. Base your drawing on the electron sea model shown in Figure 7.11.

Diagrams should show metal ions being moved into a longer, thinner form through a sea of electrons.

Everyday Chemistry

Writing in Chemistry

Sense of Danger Our sense of taste can detect certain toxins found naturally in plants. Research other modern toxins, such as lead and antifreeze, to find out why they don’t excite a negative response from our taste buds. For more on green chemistry, visit glencoe.com.

Student research should cite speculation by scientists that because humans and their ancestors did not often encounter elemental lead (or other modern toxins) in the natural environment, we have evolved having no natural aversion to eating the toxin. In fact, certain lead compounds found in paint (such as lead acetate) have a sweet taste. Research might also discuss how, contrary to the situation with lead, many plant-produced toxins have a bitter taste. This is likely due to the fact that humans and plants have co-evolved for millions of years.
Chapter 7 Assessment
pages 232–235

Section 7.1

Mastering Concepts

46. How do positive ions and negative ions form?
   An atom gains or loses electrons to achieve a stable electron configuration.

47. When do chemical bonds form?
   When a positive nucleus attracts electrons of another atom, or oppositely charged ions attract

48. Why are halogens and alkali metals likely to form ions? Explain your answer.
   Halogens need to gain only one electron to have a noble gas electron configuration. Alkali metals need to lose one.

49. The periodic table shown in Figure 7.14 contains elements labeled A–G. For each labeled element, state the number of valence electrons and identify the ion that will form.
   A: three valence electrons Al³⁺, B: two valence electrons Ba²⁺, C: one valence electron Rb⁺, D: five valence electrons N₅⁻, E: seven valence electrons I⁻, F: eight valence electrons, no ion formed, G: six valence electrons, Se²⁻

50. Discuss the importance of electron affinity and ionization energy in the formation of ions.
   High electron affinity: atom easily gains an electron; low ionization energy: atom easily loses an electron

51. The orbital notation of sulfur is shown in Figure 7.15. Explain how sulfur forms its ion.
   Sulfur gains 2 electrons in the 3p sublevel, forming a complete octet.

Mastering Problems

52. Give the number of valence electrons in an atom of each element:
   a. cesium
      1
   b. rubidium
      1
   c. gallium
      3
   d. zinc
      2
   e. strontium
      2

53. Explain why noble gases are not likely to form chemical bonds.
   They already have a full, stable outer energy level.

54. Discuss the formation of the barium ion.
   Ba will lose two electrons and form Ba²⁺, which has the stable electron configuration of Xe.

55. Explain how an anion of nitrogen forms.
   N gains three electrons, forming N₃⁻, which has the stable electron configuration of Ne.

56. The more reactive an atom, the higher its potential energy. Which atom has higher potential energy, neon or fluorine? Explain.
   Fluorine, because it will easily gain one more electron to fill its outer energy level.
57. Explain how the iron atom can form both an iron 2+ ion and an iron 3+ ion.

Iron has the electron configuration [Ar]4s²3d⁶. To form the 2+ ion the iron atom loses the 4s² electrons. When forming the 3+ ion the iron atom loses the 4s² electrons and one 3d electron.

58. Predict the reactivity of each atom based on its electron configuration.

a. potassium

Potassium, [Ar]4s¹, is reactive. It will tend to lose one outer electron and form a 1+ ion.

b. fluorine

Fluorine, [He]2s²2p⁵, is reactive. It will tend to gain one more electron and form a 1– ion.

c. neon

Neon, 1s²2s²2p⁶, will not react because it already has eight electrons in its outer energy level.

59. Discuss the formation of a 3+ scandium ion using its orbital notation, shown in Figure 7.16.

Scandium, [Ar]4s²3d¹, loses both the 4s² electrons and the 3d¹ electron to form the 3+ ion in the stable [Ar] configuration.

Section 7.2

Mastering Concepts

60. What does the term electrically neutral mean when discussing ionic compounds?

The number of electrons lost is equal to the number of electrons gained.

61. Discuss the formation of ionic bonds.

A positive ion is attracted to a negative ion and lattice energy is released.

62. Explain why potassium does not bond with neon to form a compound.

Neon already has an octet in its outer energy level; it is already stable.

63. Briefly discuss three physical properties of ionic solids that are linked to ionic bonds.

Ionic solids exist as crystals, and have high melting and boiling points.

64. Describe an ionic crystal, and explain why ionic crystals for different compounds might vary in shape.

An ionic crystal is the geometric arrangement of positive and negative ions. They vary in shape because of the size of the ions that are bonded as well as the number of each type of ion bonded together.

65. How does lattice energy change with a change in the size of an ion?

As the size on an ion increases, the lattice energy decreases.

66. In Figure 7.14, the element labeled B is barium, and the element labeled E is iodine. Explain why the compound formed between these elements will not be BaI.

Barium has two valence electrons and forms a 2+ ion. Iodine has seven valence electrons and forms a 1– ion. To form a neutral compound one Barium ion must bond with two iodide ions. A compound with one barium ion and one iodide ion is not electrically neutral and will not form.
Mastering Problems

67. Determine the ratio of cations to anions in each.
   a. potassium chloride, a salt substitute
      1:1
   b. calcium fluoride, used in the steel industry
      1:2
   c. calcium oxide, used to remove sulfur dioxide from power-plant exhaust
      1:1
   d. strontium chloride, used in fireworks
      1:2

68. Look at Figure 7.14; describe the ionic compound that forms from the elements represented by C and D.

C represents the element Rb which will lose one valence electron forming a $1^+$ ion. D represents the element nitrogen which will gain three valence electrons forming a $3^-$ ion. It will take three rubidium ions to attract to one nitride ion forming the ionic compound $\text{Rb}_3\text{N}$.

69. Discuss the formation of an ionic bond between zinc and oxygen.
   Zn will lose its outer 4s electrons, forming $\text{Zn}^{2+}$.
   Oxygen will gain the two electrons forming $\text{O}^{2-}$. $\text{Zn}^{2+}$ attracts $\text{O}^{2-}$, forming $\text{ZnO}$.

70. Using orbital notation, diagram the formation of an ionic bond between aluminum and fluorine.

71. Using electron configurations, diagram the formation of an ionic bond between barium and nitrogen.

   \[
   \text{Ba} + \text{N} \rightarrow \text{Ba}^2+ + 2\text{N}^2- + 2\text{e}^- 
   \]

   \[
   \text{[Xe]}6s^2 + \text{[He]}2s^22p^3 \rightarrow \text{[Xe]}^+ + \text{[Ne]}^2- + 2\text{e}^- 
   \]

   Barium has two valence electrons and nitrogen has five. To form a compound, a total of six electrons must transfer from three barium atoms to two nitrogen atoms.

72. Conductors Under certain conditions, ionic compounds conduct an electric current. Describe these conditions, and explain why ionic compounds are not always used as conductors.

   Ionic compounds will conduct in the molten state or dissolved in water but are nonconducting solids at room temperature.
73. Which compounds are not likely to occur: CaKr, Na₂S, BaCl₃, MgF? Explain your choices.

CaKr, because Kr is a noble gas; BaCl₃ and MgF, because charges are not balanced

74. Use Table 7.6 to determine which ionic compound has the highest melting point: MgO, KI, or AgCl. Explain your answer.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Lattice Energy</th>
<th>Compound</th>
<th>Lattice Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>KI</td>
<td>632</td>
<td>KF</td>
<td>808</td>
</tr>
<tr>
<td>KBr</td>
<td>671</td>
<td>AgCl</td>
<td>910</td>
</tr>
<tr>
<td>Rbf</td>
<td>774</td>
<td>NaF</td>
<td>910</td>
</tr>
<tr>
<td>NaI</td>
<td>682</td>
<td>LiF</td>
<td>1030</td>
</tr>
<tr>
<td>NaBr</td>
<td>732</td>
<td>SrCl₂</td>
<td>2142</td>
</tr>
<tr>
<td>NaCl</td>
<td>769</td>
<td>MgO</td>
<td>3795</td>
</tr>
</tbody>
</table>

MgO. It has the highest lattice energy and therefore requires the most energy to break the ionic bonds.

75. Which has the more negative lattice energy, CsCl or KCl? K₂O or CaO? Explain your choices.

KCl will have the more negative lattice energy because potassium is smaller than cesium and they both have a 1⁺ charge. The smaller the ion, the more negative the lattice energy. CaO will have the more negative lattice energy because the charge on the calcium ion is 2⁺ while that on potassium is 1⁺. The greater an ion's charge is, the more negative its lattice energy.

78. Discuss how an ionic compound is named.

1. Name the cation first and the anion second.
2. Monatomic cations use the element name.
3. Monatomic anions take their name from the root of the element name plus the suffix -ide.
4. Group 1 and group 2 metals have only one oxidation number. Transition metals and metals on the right side of the periodic table often have more than one oxidation number.
5. If the compound contains a polyatomic ion, simply name the ion.

79. Using oxidation numbers, explain why the formula NaF₂ is incorrect.

The oxidation number for a sodium ion is 1⁺ and that of a fluoride ion is 1⁻. The oxidation numbers indicate that one electron is lost by sodium and one electron is gained by the fluoride ion. The ions should be in a one to one ratio. The correct formula is NaF.

80. Explain what the name scandium(III) oxide means in terms of electrons lost and gained, and identify the correct formula.

Scandium(III) indicates that the scandium atom lost three electrons and oxide indicates that the oxygen atom gained two electrons. The correct formula is Sc₃O₅.

**Mastering Problems**

81. Give the formula for each ionic compound.

- calcium iodide
  \[ \text{CaI}_2 \]
- silver(I) bromide
  \[ \text{AgBr} \]
- copper(II) chloride
  \[ \text{CuCl}_2 \]
- potassium periodate
  \[ \text{KIO}_4 \]
- silver(I) acetate
  \[ \text{AgC}_2\text{H}_3\text{O}_2 \]
82. Name each of the following ionic compounds.
   a. \( \text{K}_2\text{O} \)
      potassium oxide
   b. \( \text{CaCl}_2 \)
      calcium chloride
   c. \( \text{Mg}_3\text{N}_2 \)
      magnesium nitride
   d. \( \text{NaClO} \)
      sodium hypochlorite
   e. \( \text{KNO}_3 \)
      potassium nitrate

83. Complete Table 7.14 by placing the symbols, formulas, and names in the blanks.

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{NH}_4^+ )</td>
<td>( \text{SO}_4^{2-} )</td>
<td>ammonium sulfate</td>
<td>( \text{(NH}_4\text{)}_2\text{SO}_4 )</td>
</tr>
<tr>
<td>( \text{Pb}^{2+} )</td>
<td>( \text{F}^- )</td>
<td>lead(II) fluoride</td>
<td>( \text{PbF}_2 )</td>
</tr>
<tr>
<td>( \text{Li}^+ )</td>
<td>( \text{Br}^- )</td>
<td>lithium bromide</td>
<td>( \text{LiBr} )</td>
</tr>
<tr>
<td>( \text{Na}^+ )</td>
<td>( \text{CO}_3^{2-} )</td>
<td>sodium carbonate</td>
<td>( \text{Na}_2\text{CO}_3 )</td>
</tr>
<tr>
<td>( \text{Mg}^{2+} )</td>
<td>( \text{PO}_4^{3-} )</td>
<td>magnesium phosphate</td>
<td>( \text{Mg}_3\text{(PO}_4\text{)}_2 )</td>
</tr>
</tbody>
</table>

84. Chrome Chromium, a transition metal used in chrome plating, forms both the \( \text{Cr}^{2+} \) and \( \text{Cr}^{3+} \) ions. Write the formulas for the ionic compounds formed when each of these ions react with fluorine and oxygen ions.
   - fluorine: \( \text{CrF}_2, \text{CrF}_3 \); oxygen: \( \text{CrO}, \text{Cr}_2\text{O}_3 \)

85. Which are correct formulas for ionic compounds? For those that are not correct, give the correct formula and justify your answer.
   a. \( \text{AlCl}_3 \)
      \( \text{AlCl}_3 \); one \( \text{Al}^{3+} \) ion bonds to three \( \text{Cl}^- \) ions
   b. \( \text{Na}_2\text{SO}_4 \)
      \( \text{Na}_2\text{SO}_4 \); two \( \text{Na}^+ \) ions bond to \( \text{SO}_4^{2-} \)
   c. \( \text{BaOH}_2 \)
      \( \text{Ba(OH)}_2 \); if a polyatomic ion needs a subscript, parentheses must be used
   d. \( \text{Fe}_2\text{O}_3 \)
      \( \text{Fe}_2\text{O}_3 \) or \( \text{FeO} \); iron forms either \( \text{Fe}^{2+} \) or \( \text{Fe}^{3+} \)

86. Write the formulas for all of the ionic compounds that can be formed by combining each of the cations with each of the anions listed in Table 7.15. Name each compound formed.
   - \( \text{K}_2\text{SO}_3 \), potassium sulfite; \( \text{KI} \), potassium iodide; \( \text{KNO}_3 \), potassium nitrate; \( \text{(NH}_4\text{)}_2\text{SO}_3 \), ammonium sulfite; \( \text{NH}_4\text{I} \), ammonium iodide; \( \text{NH}_4\text{NO}_3 \), ammonium nitrate; \( \text{Fe}_2\text{(SO}_3\text{)}_3 \), iron(III) sulfite; \( \text{Fe}_3\text{I}_3 \), iron(III) iodide; \( \text{Fe(NO}_3\text{)}_3 \), iron(III) nitrate

Section 7.4

Mastering Concepts

87. Describe a metallic bond.
   Each positive metal ion is attracted to delocalized valence electrons.

88. Briefly explain why metallic alloys are made.
   Alloys have slightly different properties than those of the pure metal they are mixed from. Some alloys have increased strength and are harder than the pure metal.

89. Briefly describe how malleability and ductility of metals are explained by metallic bonding.
   When a force is applied to a metallic solid, the metal ions move as well as the delocalized electrons.
90. Compare and contrast the two types of metal alloys.

A substitutional alloy has metal atoms that are similar in size. An interstitial alloy has two differently sized atoms.

91. Explain how a metallic bond is similar to an ionic bond.

The bonds are similar because they are both formed by the attraction of oppositely charged particles.

92. Brass Copper and zinc are used to form brass, an alloy. Briefly explain why these two metals form a substitutional alloy and not an interstitial alloy.

The copper and zinc ions are about the same size. They can replace each other in the metal forming a substitutional alloy.

93. How is a metallic bond different from an ionic bond?

A metallic bond is an electrostatic attraction between a positive metal ion and a “sea” of free valence electrons; an ionic bond is an electrostatic attraction between a positive metallic ion and a negative nonmetallic ion.

94. Silver Briefly explain why silver is a good conductor of electricity.

It has delocalized electrons that are free to move.

95. Steel Briefly explain why steel, an alloy of iron, is used to build the supporting structure of many buildings.

Iron forms a strong metallic bond, giving solid iron hardness and strength.

96. The melting point of beryllium is 1287°C, while that of lithium is 180°C. Explain the large difference in values.

Beryllium has two delocalized electrons per atom. Lithium has one. As the number of delocalized electrons increases, lattice energy increases raising the melting point.

97. Titanium has a boiling point of 3287°C and copper has a boiling point of 2567°C. Explain why there is a difference in the boiling points of these two metals.

Titanium can have up to four delocalized electrons and copper has up to two. The metallic bonding in titanium is greater than that in copper.

98. Alloys Describe the difference between the metal alloy sterling silver and carbon steel in terms of the types of alloys involved.

Sterling silver is substitutional, formed from silver and copper. Carbon steel is interstitial, formed from iron and carbon.

Mixed Review

99. Give the number of valence electrons for atoms of oxygen, sulfur, arsenic, phosphorus, and bromine.

6, 6, 5, 5, and 7, respectively.

100. Explain why calcium can form a Ca^{2+} ion but not a Ca^{3+} ion.

Ca, [Ar]4s^2, will lose two electrons. If it loses an inner 3p electron, it is unstable.

101. Which ionic compounds would have the most negative lattice energy: NaCl, KCl, or MgCl_2? Explain your answer.

MgCl_2, lattice energy increases with increased charge.

102. Give the formula for each ionic compound.

a. sodium sulfide
   \( \text{Na}_2\text{S} \)

b. iron(III) chloride
   \( \text{FeCl}_3 \)

c. sodium sulfate
   \( \text{Na}_2\text{SO}_4 \)

d. calcium phosphate
   \( \text{Ca}_3(\text{PO}_4)_2 \)

e. zinc nitrate
   \( \text{Zn(NO}_3)_2 \)
103. Cobalt, a transition metal, forms both the $\text{Co}^{2+}$ and $\text{Co}^{3+}$ ions. Write the correct formulas, and give the name for the oxides formed by the two different ions.

CoO, cobalt(II) oxide; $\text{Co}_2\text{O}_3$, cobalt(III) oxide

104. Complete Table 7.16.

<table>
<thead>
<tr>
<th>Element</th>
<th>Valence Electrons</th>
<th>Ion Formed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Selenium</td>
<td>6</td>
<td>$\text{Se}^{2-}$</td>
</tr>
<tr>
<td>Tin</td>
<td>4</td>
<td>$\text{Sn}^{2+}$</td>
</tr>
<tr>
<td>Iodine</td>
<td>7</td>
<td>$\text{I}^-$</td>
</tr>
<tr>
<td>Argon</td>
<td>8</td>
<td>none</td>
</tr>
</tbody>
</table>

105. Gold Briefly explain why gold can be used both in jewelry and as a conductor in electronic devices.

Delocalized electrons allow it to conduct. It is malleable and ductile.

106. Discuss the formation of the nickel ion with a 2+ oxidation number.

Nickel, [Ar]3d$^8$4s$^2$, will lose the two outer 4s electrons.

107. Compare the oxyanions sulfate and sulfite.

Both oxyanions are composed of sulfur and oxygen. Sulfate $\text{SO}_4^-$ has more oxygen atoms than sulfite, $\text{SO}_3^-$. The prefix $\text{-ite}$ is used to indicate one less oxygen atom. Both ions have the same oxidation number of 2$\text{−}$.

108. Using electron-dot structures, diagram the formation of an ionic bond between potassium and iodine.

$\text{K}^+ + \text{I}^- \rightarrow [\text{K}]^+ + [\text{I}]^-$

109. Magnesium forms both an oxide and a nitride when burned in air. Discuss the formation of magnesium oxide and magnesium nitride when magnesium atoms react with oxygen and nitrogen atoms.

A Mg atom loses two valence electrons, forming $\text{Mg}^{2+}$. An O atom gains two electrons, forming $\text{O}^{2-}$. One magnesium ion attracts one oxygen ion, forming $\text{MgO}$. Three Mg atoms each lose two electrons, forming three $\text{Mg}^{2+}$, while two N atoms each gain three electrons, forming two $\text{N}^3-$. The ions attract each other, forming $\text{Mg}_3\text{N}_2$.

110. An external force easily deforms sodium metal, while sodium chloride shatters when the same amount of force is applied. Why do these two solids behave so differently?

Sodium metal contains metallic bonds. Sodium chloride is an ionic solid.

111. Name each ionic compound.

a. CaO calcium oxide
b. BaS barium sulfide
c. AlPO$_4$ aluminum phosphate
d. Ba(OH)$_2$ barium hydroxide
e. Sr(NO$_3$)$_2$ strontium nitrate

112. Design a concept map to explain the physical properties of both ionic compounds and metallic solids.

Concept maps will vary.
113. **Predict** which solid in each pair will have the higher melting point. Explain your answers.

   a. NaCl or CsCl
      NaCl; smaller ion size

   b. Ag or Cu
      Cu; it is smaller

   c. Na₂O or MgO
      MgO; Mg has a greater charge

114. **Compare and contrast** cations and anions.

   Cations are formed from the loss of electrons and have a positive charge. Anions are formed from the gain of electrons and have a negative charge.

115. **Observe and Infer** Identify the mistakes in the incorrect formulas and formula names, and design a flow chart to prevent the mistakes.

   a. copper acetate
      Identify the metal as copper(II) or copper(I).

   b. Mg₂O₂
      The formula unit is not in the simplest ratio.

   c. Pb₂O₅
      Lead can have only the oxidation state of 2⁺ or 4⁺, not 5⁺.

   d. disodium oxide
      Prefixes are not used in ionic compounds.

   e. Al₂SO₄₃
      If a polyatomic ion requires a subscript, use parentheses.

   Flow charts will vary.

116. **Apply Concepts** Examine the ions in the beaker shown in Figure 7.17. Identify two compounds that could form using the available ions, and explain why this is possible.

   The compounds that can be formed are: Al₂S₃, AlN, AlF₃, Na₂S, Na₂N, NaF, CaS, Ca₃N₂₁, and CaF₂. Students should explain about electrons transferred from the atom to form a positive ion as well as electrons gained by atoms to form negative ions. They should also discuss the attraction between positive ions and negative ions and the forming of an electrically neutral compound.

117. **Apply** Praseodymium is a lanthanide element that reacts with hydrochloric acid, forming praseodymium(III) chloride. It also reacts with nitric acid, forming praseodymium(III) nitrate. Praseodymium has the electron configuration [Xe]4f³6s².

   a. Examine the electron configuration, and explain how praseodymium forms a 3⁺ ion.
      Praeseodymium must lose the outer 6s² electrons and one of the 4f electrons to form the 3⁺ ion.

   b. Write the correct formulas for both compounds formed by praseodymium.
      The compounds formed are PrCl₃ and Pr(NO₃)₃.

118. **Hypothesize** Look at the locations of potassium and calcium on the periodic table. Form a hypothesis to explain why the melting point of calcium is considerably higher than the melting point of potassium.

   Calcium has two delocalized electrons for every one for potassium. Thus, calcium has a higher melting point.

119. **Assess** Explain why the term *delocalized* is an appropriate term for the electrons involved in metallic bonding.

   They are free to move; the electrons are not held to any specific atom.
120. **Apply** All uncharged atoms have valence electrons. Explain why elements such as iodine and sulfur do not have metallic bonds. They gain electrons. Thus, their electrons are not delocalized.

121. **Analyze** Explain why lattice energy is a negative quantity. Lattice energy is the energy released when an ionic bond forms. Therefore, bonded ions have lower potential energy than separated ions. By convention, the separated state is considered the reference; therefore, the bonded state will have a negative value.

### Challenge Problem

122. **Ionic Compounds** Chrysoberyl is a transparent or translucent mineral that is sometimes opalescent. It is composed of beryllium aluminum oxide, BeAl₂O₄. Identify the oxidation numbers of each of the ions found in this compound. Explain the formation of this ionic compound.

Be, a Group 2 element, forms a 2⁺ ion; Al, a Group 13 element, forms a 3⁺ ion; and O, a Group 16 element, forms a 2⁻ ion. There are two electrons lost from one atom of beryllium and six electrons lost from two atoms of aluminum. Four oxygen atoms gain a total of eight electrons. The positive ions attract the negative ions forming an electrically neutral compound.

### Cumulative Review

123. You are given a liquid of unknown density. The mass of a graduated cylinder containing 2.00 mL of the liquid is 34.68 g. The mass of the empty graduated cylinder is 30.00 g. Given this information, determine the density of the liquid. (Chapter 2)

mass of the liquid = 34.68 g − 30.00 g = 4.68 g

\[
\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.68 \text{ g}}{2.00 \text{ mL}} = 2.34 \text{ g/mL}
\]

124. In the laboratory, students used a balance and a graduated cylinder to collect the data shown in Table 7.17. Calculate the density of the sample. If the accepted value of this sample is 7.01 g/mL, calculate the percent error.

(Chapter 2)

<table>
<thead>
<tr>
<th>Volume and Mass Data</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of sample</td>
</tr>
<tr>
<td>Volume of water alone</td>
</tr>
<tr>
<td>Volume of water + sample</td>
</tr>
</tbody>
</table>

\[
\text{volume of sample} = 43.1 \text{ mL} - 39.0 \text{ mL} = 4.1 \text{ mL}
\]

\[
\text{mass of sample} = 19.21 \text{ g}
\]

\[
\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{19.21 \text{ g}}{4.1 \text{ mL}} = 4.7 \text{ g/mL}
\]

\[
\text{percent error} = \left( \frac{(4.7 \text{ g/mL}) - (7.01 \text{ g/mL})}{7.01 \text{ g/mL}} \right) 
= -33\%
\]

The students calculated density is 33% too low.

125. A mercury atom drops in energy from \(1.413 \times 10^{-18} \text{ J}\) to \(1.069 \times 10^{-18} \text{ J}\). (Chapter 5)

a. What is the energy of the photon emitted by the mercury atom?

\[\Delta E = 1.413 \times 10^{-18} \text{ J} - 1.069 \times 10^{-18} \text{ J} = 3.44 \times 10^{-19} \text{ J}\]

b. What is the frequency of the photon emitted by the mercury atom?

\[
\nu = \frac{\Delta E}{h} = \frac{3.44 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} = 5.19 \times 10^{14} \text{ s}^{-1}
\]

c. What is the wavelength of the photon emitted by the mercury atom?

\[
\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{5.19 \times 10^{14} \text{ s}^{-1}} = 5.78 \times 10^{-7} \text{ m}
\]

or 578 nm

126. Which element has the greater ionization energy, chlorine or carbon? (Chapter 6)

chlorine
127. Compare and contrast the ways in which metals and nonmetals form ions, and explain why they are different. (Chapter 6)

Metals lose electrons to form cations; nonmetals gain electrons to form anions. Both form ions to gain stability.

128. What are transition elements? (Chapter 6)

the d block elements

129. Write the symbol and name of the element that fits each description. (Chapter 6)

a. the second-lightest of the halogens
   
   \( \text{Cl, chlorine} \)

b. the metalloid with the lowest period number

   \( \text{B, boron} \)

c. the only group 16 element that is a gas at room temperature

   \( \text{O, oxygen} \)

d. the heaviest of the noble gases

   \( \text{Rn, radon} \)

e. the group 15 nonmetal that is a solid at room temperature

   \( \text{P, phosphorus} \)

Additional Assessment

Writing in Chemistry

130. Free Radicals Many researchers believe that free radicals are responsible for the effects of aging and cancer. Research free radicals, and write about the cause and what can be done to prevent free radicals.

Student answers will vary. Students should discuss reduction and oxidation (gain and loss of electrons) in forming free radicals as well as antioxidants, Vitamin E and Vitamin C.

131. Growing Crystals Crystals of ionic compounds can be easily grown in the laboratory setting. Research the growth of crystals, and design an experiment to grow a crystal in the laboratory.

Student answers will vary. Students should include the use of supersaturated solutions and the evaporation of water from the solution allows crystals to grow large over a period of time.

Document-Based Questions

Oceans As part of an analysis of the world’s oceans, scientists summarized the ion-related data shown in Table 7.18.

Data from: Royal Society of Chemistry, *All at sea? The chemistry of the oceans.*

<table>
<thead>
<tr>
<th>Ion</th>
<th>Concentration (mg/dm³)</th>
<th>% by mass (of the total dissolved solids)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl⁻</td>
<td>19,000</td>
<td>55.04</td>
</tr>
<tr>
<td>Na⁺</td>
<td>10,500</td>
<td>30.42</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>2655</td>
<td>7.69</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>1350</td>
<td>3.91</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>400</td>
<td>1.16</td>
</tr>
<tr>
<td>K⁺</td>
<td>380</td>
<td>1.10</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>140</td>
<td>0.41</td>
</tr>
<tr>
<td>Br⁻</td>
<td>65</td>
<td>0.19</td>
</tr>
<tr>
<td>BO₃³⁻</td>
<td>20</td>
<td>0.06</td>
</tr>
<tr>
<td>SiO₃²⁻</td>
<td>8</td>
<td>0.02</td>
</tr>
<tr>
<td>Sr²⁺</td>
<td>8</td>
<td>0.02</td>
</tr>
<tr>
<td>F⁻</td>
<td>1</td>
<td>0.003</td>
</tr>
</tbody>
</table>

132. Identify the anions and cations listed in Table 7.18.

anions: chloride, Cl⁻, sulfate, SO₄²⁻, carbonate, CO₃²⁻, bromide, Br⁻, borate, BO₃³⁻, silicate, SiO₃²⁻, fluoride, F⁻

cations: sodium, Na⁺, magnesium, Mg²⁺, strontium, Sr²⁺, calcium, Ca²⁺, potassium, K⁺
133. Create a bar graph of each ion’s concentration. Explain why this is a difficult graph to draw.

Bar charts should correspond to concentration data from Table 7.18. The graph is difficult to draw because the range of the data is so great; some data are very small and some data are very large. This difficulty can be eased by using a logarithmic scale on the vertical axis.

134. Sodium chloride is not the only ionic compound that forms from sea water. Identify four other compounds that could be formed that contain the sodium ion. Write both the formula and the name for each compound.

Students may identify the following compounds: sodium chloride, NaCl; sodium sulfate, Na₂SO₄; sodium carbonate, Na₂CO₃; sodium bromide, NaBr; sodium borate, Na₃BO₃; sodium silicate, Na₂SiO₃; sodium fluoride, NaF.

Standardized Test Practice
pages 236–237

Multiple Choice

Use the figure below to answer Question 1.

1. Which description is supported by the model shown?
   a. Metals are shiny, reflective substances.
   b. Metals are excellent conductors of heat and electricity.
   c. Ionic compounds are malleable compounds.
   d. Ionic compounds are good conductors of electricity.
   b

2. Which is NOT true of the Sc³⁺ ion?
   a. It has the same electron configuration as Ar.
   b. It is a scandium ion with three positive charges.
   c. It is considered to be a different element than a neutral Sc atom.
   d. It was formed by the removal of the valence electrons of Sc.
   c

3. Of the salts below, which would require the most energy to break the ionic bonds?
   a. BaCl₂
   b. LiF
   c. NaBr
   d. KI
   a

4. The high strength of its ionic bonds results in all of the following properties of NaCl EXCEPT
   a. hard crystals
   b. high boiling point
   c. high melting point
   d. low solubility
   d

5. Which is the correct formula for the compound chromium (III) sulfate?
   a. Cr₃SO₄
   b. Cr₃(SO₄)₃
   c. Cr₂(SO₄)₂
   d. Cr(SO₄)₃
   b
Use the table below to answer Questions 6–8.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Bond Type</th>
<th>Melting Point(°C)</th>
<th>Boiling Point(°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>F₂</td>
<td>Nonpolar covalent</td>
<td>−220</td>
<td>−188</td>
</tr>
<tr>
<td>CH₄</td>
<td>Nonpolar covalent</td>
<td>−183</td>
<td>−162</td>
</tr>
<tr>
<td>NH₃</td>
<td>Polar covalent</td>
<td>−78</td>
<td>−33</td>
</tr>
<tr>
<td>CH₃Cl</td>
<td>Polar covalent</td>
<td>−64</td>
<td>61</td>
</tr>
<tr>
<td>KBR</td>
<td>Ionic</td>
<td>730</td>
<td>1435</td>
</tr>
<tr>
<td>Cr₂O₃</td>
<td>Ionic</td>
<td>?</td>
<td>4000</td>
</tr>
</tbody>
</table>

6. A compound is discovered to have a melting point of −100°C. Which could be true of this compound?
   a. It definitely has an ionic bond.
   b. It definitely has a polar covalent bond.
   c. It has either a polar covalent bond or a nonpolar covalent bond.
   d. It has either a polar covalent bond or an ionic bond.

   c

7. Which could NOT be the melting point of Cr₂O₃?
   a. 2375°C
   b. 950°C
   c. 148°C
   d. 3342°C

   c

8. Which is supported by the data in the table?
   a. Nonpolar covalent bonds have high boiling points.
   b. Polar covalent bonds have high melting points.
   c. Ionic bonds have low melting points.
   d. Ionic bonds have high boiling points.

   d

9. Which is the correct orbital diagram for the third and fourth principal energy levels of vanadium?
   a. ![Diagram a]
   b. ![Diagram b]
   c. ![Diagram c]
   d. ![Diagram d]

   a

Short Answer

Use the table below to answer Questions 10–12.

Lutetium is a rare earth element that can be used to speed up the chemical reactions involved in petroleum processing. It has two naturally occurring isotopes.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Form of Decay</th>
<th>Percent Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>¹⁷⁵Lu</td>
<td>None</td>
<td>97.41</td>
</tr>
<tr>
<td>¹⁷⁶Lu</td>
<td>Beta</td>
<td>2.59</td>
</tr>
</tbody>
</table>

10. Show the setup and calculate the average atomic mass of lutetium.

   $$M_{avg} = \frac{(0.9741 \times 175) + (0.0259 \times 176)}{2} \text{ amu}$$

   $$M_{avg} = 175.03 \text{ amu}$$; The answer should have 3 significant digits because the atomic masses only have 3 significant digits.

11. Identify the product when lutetium goes through nuclear decay.

   ¹⁷⁶⁵Hf

12. Compare the number of protons and neutrons in each of these isotopes.

   ¹⁷⁵⁷Lu: 71 protons, 104 neutrons

   ¹⁷⁶⁷Lu: 71 protons, 105 neutrons
Extended Response

13. Relate the change in atomic radius to the changes in atomic structure that occur across the periodic table.

Atomic radii generally decrease across a given period because the increasing positive charge in the nucleus tends to contract the outer electron orbitals that are being filled at the same principal energy level.

Atomic radii generally increase down a given group because new principal energy levels (at larger radii) are being added to the atom. Increasing positive charge in the nucleus is not sufficient to overcome this effect.

Use the diagram below to answer Question 14.

14. Relate the change in ionic radius to the changes in ion formation that occur across the periodic table.

Cations are formed from their corresponding neutral atoms by releasing outer valence electrons to achieve a stable noble gas configuration. The ionic radius is smaller than the neutral atomic radius because all of the valence electrons comprising the highest principal energy level are being released.

SAT Subject Test: Chemistry

Use the diagram below to answer Question 15.

15. Which describes the state of matter shown?
   a. solid, because the particles are tightly packed against one another
   b. gas, because the particles are flowing past one another
   c. liquid, because the particles are able to move freely
   d. solid, because there is a regular pattern to the particles
   e. liquid, because the particles are flowing past one another
e

Use the list of elements below to answer Questions 16–20.

a. sodium  
   b. chromium  
   c. boron  
   d. argon  
   e. chlorine

16. Which has its outermost electrons in an s-sublevel?
   a.

17. Which has seven valence electrons?
   e.

18. Which is a transition metal?
   c.

19. Which has an electron configuration of 1s^22s^22p^63s^23p^5?
   e.

20. Which is a noble gas?
   d.