

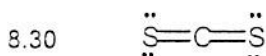
- (c) Se^{3+} : $[\text{Ar}]4s^23d^{10}4p^1$ (This is a very unlikely ion. A more stable and commonly-found ion would be Sc^{3+} .) Sc^{3+} : $[\text{Ar}]$, noble-gas configuration
- (d) Ru^{2+} : $[\text{Kr}]4d^6$ (e) Tl^+ : $[\text{Xe}]6s^24f^{14}5d^{10}$ (f) Au^+ : $[\text{Xe}]4f^{14}5d^{10}$

D.S.
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8.17 KF , 808 kJ/mol; CaO , 3414 kJ/mol; ScN , 7547 kJ/mol

The sizes of the ions vary as follows: $\text{Sc}^{3+} < \text{Ca}^{2+} < \text{K}^+$ and $\text{F}^- < \text{O}^{2-} < \text{N}^{3-}$. Therefore, the interionic distances are similar. According to Coulomb's law for compounds with similar ionic separations, the lattice energies should be related as the product of the charges of the ions. The lattice energies above are approximately related as (1)(1): (2)(2): (3)(3) or 1 : 4 : 9. Slight variations are due to the small differences in ionic separations.

- 8.20 (a) In MgO , the magnitude of the charges on both ions is 2; in MgCl_2 , the magnitudes of the charges are 2 and 1. Also, the Cl^- ion is larger than the O^{2-} ion, so the charge separation is greater in MgCl_2 . Thus, the lattice energy of MgO is greater, because the product of the ionic charges is greater and the ion separation is smaller.
- (b) The ions have +1 and -1 charge in all three compounds. In NaCl the cationic and anionic radii are smaller than in the other two compounds, so it has the largest lattice energy. In RbBr and CsBr , the anion is the same, but the Cs cation is larger, so CsBr has the smaller lattice energy.



The C-S bonds in CS_2 are double bonds, so the C-S distances will be shorter than a C-S single bond distance.

8.34 Electronegativity increases going up and to the right in the periodic table.

- (a) O (b) Al (c) Cl (d) F

8.36 The more different the electronegativity values of the two elements, the more polar the bond.

- (a) $\text{O-F} < \text{C-F} < \text{Be-F}$. This order is clear from the periodic trend.
- (b) $\text{Br-N} < \text{P-Br} < \text{Br-O}$. Refer to the electronegativity values in Figure 8.9 to confirm the order of bond polarity. The electronegativity order of the four elements is $\text{P} < \text{Br} < \text{N} < \text{O}$. Br is significantly to the right and below the other three elements, but these directions have conflicting trends. Even though Br and N are farthest apart, their electronegativity values are nearly equal. The Br-N bond is least polar, with Br the less electronegative element.
- (c) $\text{C-S} < \text{N-O} < \text{B-F}$. You might predict that N-O is least polar since the elements are adjacent on the table. However, the big decrease going from the second row to the third means that the electronegativity of S is not only less than that of O, but essentially the same as that of C. C-S is the least polar.

8.42 Generally, compounds formed by a metal and a nonmetal are described as ionic, while compounds formed from two or more nonmetals are covalent.

- (a) MnF_3 , ionic (b) CrO_3 , ionic
- (c) AsBr_5 , ionic (Although As is a metalloid, use of a roman numeral usually presumes an ionic compound.)
- (d) sulfur tetrafluoride, covalent (e) molybdenum(IV) chloride, ionic

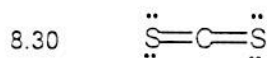
- 8.14 (a) Zn^{2+} : $[\text{Ar}]3d^{10}$ (b) Te^{4-} : $[\text{Kr}]5s^2 4d^{10} 5p^6 = [\text{Xe}]$, noble-gas configuration
 (c) Se^{3+} : $[\text{Ar}]4s^2 3d^{10} 4p^1$ (This is a very unlikely ion. A more stable and commonly-found ion would be Sc^{3+} .) Sc^{3+} : $[\text{Ar}]$, noble-gas configuration
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The C-S bonds in CS_2 are double bonds, so the C-S distances will be shorter than a C-S single bond distance.

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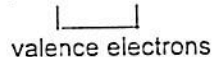
- 8.36 The more different the electronegativity values of the two elements, the more polar the bond.

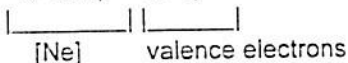
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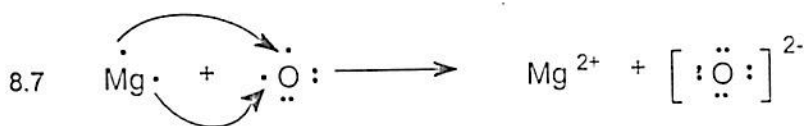
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 (d) sulfur tetrafluoride, covalent (e) molybdenum(IV) chloride, ionic

8.1 (a) Valence electrons are those that take part in chemical bonding, those in the outermost electron shell of the atom. This usually means the electrons beyond the core noble-gas configuration of the atom, although it is sometimes only the outer shell electrons.

(b) N: [He] $2s^2 2p^3$ A nitrogen atom has 5 valence electrons.


(c) $1s^2 2s^2 2p^6 3s^2 3p^2$ The atom (Si) has 4 valence electrons.




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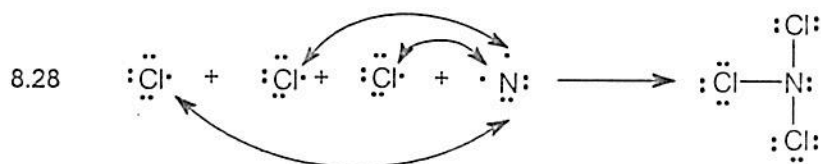
8.11 (a) AlF_3 (b) K_2S (c) Y_2O_3 (d) Mg_3N_2

8.25 (a) A *covalent bond* is the bond formed when two atoms share one or more pairs of electrons.

(b) The ionic bonding in $NaCl$ is due to strong electrostatic attraction between oppositely charged Na^+ and Cl^- ions. The covalent bonding in Cl_2 is due to sharing of a pair of electrons by two neutral chlorine atoms.

8.26 K and Ar. K is an active metal with one valence electron. It is most likely to achieve an octet by losing this single electron and to participate in ionic bonding. Ar has a stable octet of valence electrons; it is not likely to form chemical bonds of any type.

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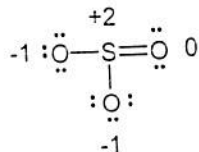
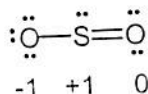
8.32 (a) The electronegativity of the elements increases going from left to right across a row of the periodic chart.

(b) Electronegativity decreases going down a family of the periodic chart.

(c) Generally, the trends in electronegativity are the same as those in ionization energy and opposite those in electron affinity. That is, the more positive the ionization energy and the more negative the electron affinity (omitting a few exceptions), the greater the electronegativity of an element.

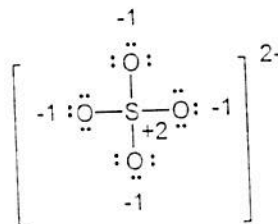
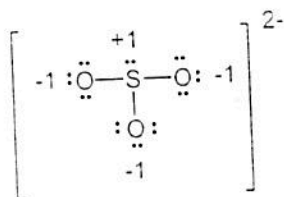
Check. Each atom is surrounded by an octet (or duet) and the sum of the formal charges in particle is the charge on the particle.

- 8.46 (a) 18 e⁻, 9 e⁻ pairs (b) 24 e⁻, 12 e⁻ pairs



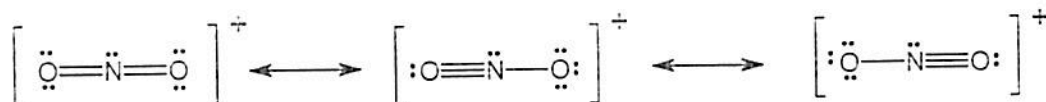
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- (c) 26 e⁻, 13 e⁻ pairs (d) 32 e⁻, 16 e⁻ pairs



- (c) Since each N–O bond has partial double bond character, the N–O bond length in NO₂⁻ should be shorter than in species with formal N–O single bonds.

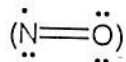
- 8.48 (a) 16 e⁻, 8 e⁻ pairs



- (b) More than one correct Lewis structure can be drawn, so resonance structures are needed to accurately describe the structure.
- (c) NO₂⁺ has 16 valence electrons. Consider other triatomic molecules involving second-row nonmetallic elements. O₃²⁺ or C₃⁴⁺ are not 'common' (or stable). CO₂ is common and matches the description (as does N₃⁻, azide ion).

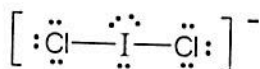
- 8.56 In the third period, atoms have the space and available orbitals to accommodate extra electrons. Since atomic radius increases going down a family, elements in the third period and beyond are less subject to destabilization from additional electron-electron repulsions. Also, the third shell contains d orbitals that are relatively close in energy to 3s and 3p orbitals (the ones that accommodate the octet) and provide an allowed energy state for the extra electrons.

- 8.58 (a) 11 e⁻, 5.5 e⁻ pairs;
odd electron molecule



The odd electron is probably on N
because it is less electronegative than O.

- (b) 22 e⁻, 11 e⁻ pairs



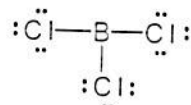
10 e⁻ around central I

- (c) 18 e⁻, 9 e⁻ pairs



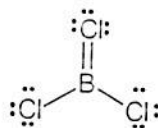
Structures with expanded octets
around S can be drawn, but are
not preferred (Section 8.7).

- (d) 24 e⁻, 12 e⁻ pairs

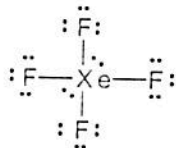


6 e⁻ around B

Structures such as the one below
can be drawn, but are not likely
because of high formal charges on
B and Cl.



- (e) 36 e⁻, 18 e⁻ pairs



12 e⁻ around the central Xe

(F cannot expand beyond an octet.)

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