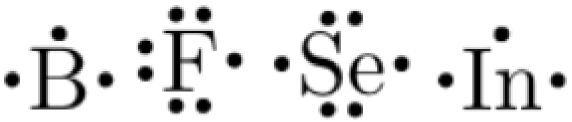
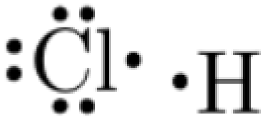
7.7-7.14, 7.16-7.17, 7.19-7.29, 7.31-7.40

7.7 An Na+ ion is not found in nature, because it involves removing a core electron. There is always a large jump in ionization energy whenever removing an electron disrupts an electron configuration ending in np6.
7.8 Nonmetals tend to form anions rather than cations because atoms that gain electrons most readily are also ones that would need to lose large numbers of electrons to produce cations with stable np6 electron configurations.
7.9 (A) N; (B) Ba2+; (C) Se; (D) Co3+
7.10 (A) Ga3+ < Ca2+ < K+ ; (B) Be2+ < Mg2+ < Ca2+ < Ba2+; (C) Al3+ < K+ < Sr2+ < Rb+; (D) Ca2+ < K+ < Rb+
PART A:
These are isoelectronic species, which means that they all have the same number of electrons. However, they have different numbers of protons. Having fewer protons results in a lesser pull on the electrons, and therefore a larger radius. Therefore, a decrease in the atomic number determines the order of increasing ionic radii. Ga3+ < Ca2+ < K+
PART C:
Al3+ has the lowest principal quantum number, so it is expected to have the smallest radius. K+ has the second lowest principal quantum number, so it is expected to have the second smallest radius. Rb+ and Sr2+ have the highest principal quantum number. These are isoelectronic species, which means that they all have the same number of electrons. Sr2+ has the third smallest radius because it is located to the right of Rb+ in the same row. Therefore, Rb+ must have the largest ionic radius. Al3+<K+<Sr2+<Rb+
7.11 (A) Cl- < S2- < P3-; (B) O2- < S2- < Se2-; (C) N3- < S2- < Br-; (D) Cl- < Br- < I-
7.12 The product of charges for the Na/F pair is (1+)(1-) = 1-, whereas the product of charges for the Mg/F pair is (2+)(1-) = 2-. Since according to Coulomb's law, lattice energy is directly proportional to the magnitude of the product of two charges, then we conclude that the Mg/F pair will have the larger lattice energy.
7.13 PART A:
Charges alternate in sign in a lattice. Therefore, first nearest neighbors exhibit acttractive forces, second nearest neighbors exhibit repulsive forces, third nearest neighbors exhibit attractive forces, and finally fourth nearest neighbors exhibit repulsive forces.
PART B:
With each successive neighbor, the distance between charges increases. According to Coulomb's law, potential energy is inversely proportional to the distance between two charges. So if the distance between them increases, then the magnitude of the potential energy between them must decrease. Therefore, ｜V1st｜>｜V2nd｜>｜V3rd｜>｜V4th｜.
7.14 An ionic bond involves the transfer of electrons from one atom to another, whereas a covalent bond involves the sharing of electrons between pairs of atoms.
7.16 As atoms begin to come closer together, the electrons from one atom become attracted to the positively charged nucleus of another atom. This attraction lowers the

potential of the system.
7.17 Coulombic forces are involved in covalent bonding, because the shared pairs of electrons are attracted to the protons from the nuclei of both atoms involved in the bonding. These shared pairs also experience repulsive forces from other electrons.
7.19 Atoms of second period elements cannot have more than eight electrons around them due to the lack of existence of 2d orbitals. However, elements in the third period and beyond have nd orbitals available to form expanded octets. Since electron pairs take up space and repel other electron pairs surrounding the central atom, then there is a limit to how many bonds can form around the central atom.

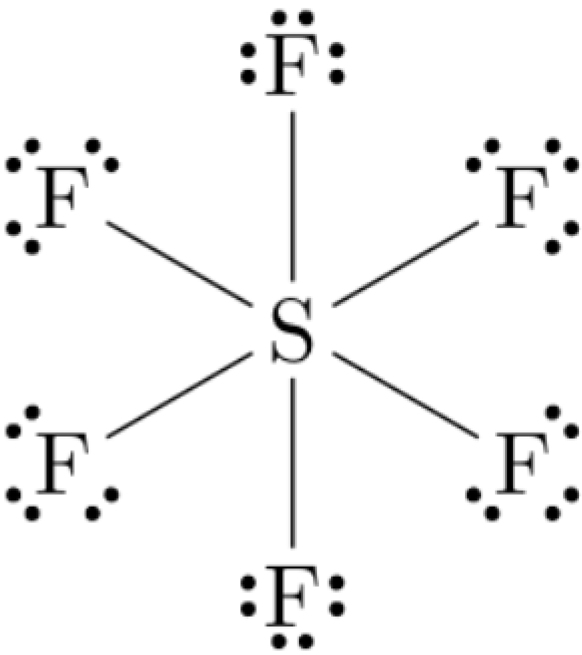
7.20 

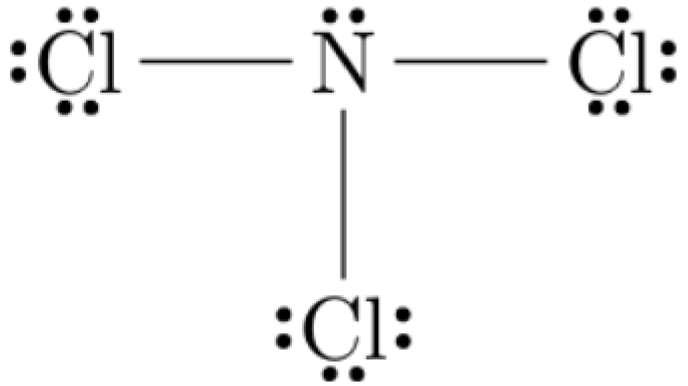
7.21 

7.22  Both chlorine and hydrogen each have a single unpaired electron. When these two atoms combine to form a compound, these unpaired electrons combine to form a shared paired of electrons. The octet rule is satisfied and additional hydrogen atoms cannot be bonded to chlorine.
7.23 A lone pair is a pair of electrons associated with only a single atom.
7.24 PART A: two electrons; PART B: four electrons; PART C: six electrons
7.25 The bond energy of a double bond is higher than the bond energies of any of the single bonds, but not twice as high. This explains the energetically favored process in which a double bond is transformed into two single bonds.
7.26 Electronegativity is a relative measure of the tendency of an atom in a molecule to attract the shared electrons in a covalent bond.
7.27 Electron affinity is the attraction of an atom for an electron, whereas electronegativity is the relative attraction of an atom for a shared pair of electrons in a covalent bond.
7.28 Noble gases have a completely filled outer shell of valence electrons and have no tendency to form molecules with covalent bonds. Many elements in the seventh period are radioactive with very short half-lives.
7.29 Covalent bonding between atoms of different electronegativity results in a bond in which there is a separation of positive and negative charge centers. The more electronegative atom becomes a center of partial negative charge because it has more electron density than is necessary to balance its own nuclear charge. The more electropositive atom takes on a partial positive charge, because it has insufficient electron density to balance its own nuclear charge.
7.31 A bond between two atoms with different electronegativities is called a polar bond, because of the dipole moment resulting from the asymmetrical charge distribution

caused by the unequal sharing of electron pairs in a covalent bond.
7.32 PART A: ionic bond; PART B: covalent bond; PART C: covalent bond; PART D: ionic bond; PART E: covalent bond; PART F: ionic bond
PART C:
Br and I are both nonmetals. In general, the bond between two nonmetals is a covalent bond. Therefore, we conclude that Br and I forms a covalent bond.
7.33 PART A: O−H has the most polar bond. C−H and S−H have the least polar bond.
PART B: H−Cl has the most polar bond. Cl−Cl has the least polar bond.
PART C: C−F has the most polar bond. F−F has the least polar bond.
PART D: N−H has the most polar bond. N−Cl has the least polar bond.
7.34 Electronegativity values for metals range in value from 0.8 to 1.9, which corresponds to an electronegativity difference of 1.1. Generally speaking, an electronegativity difference of at least 2.0 between two atoms is required for an ionic bond to form.
7.35 A
7.36 BrCl

7.37  





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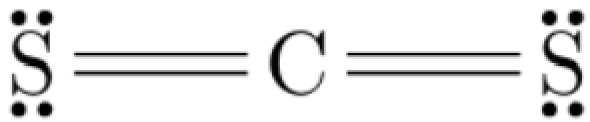
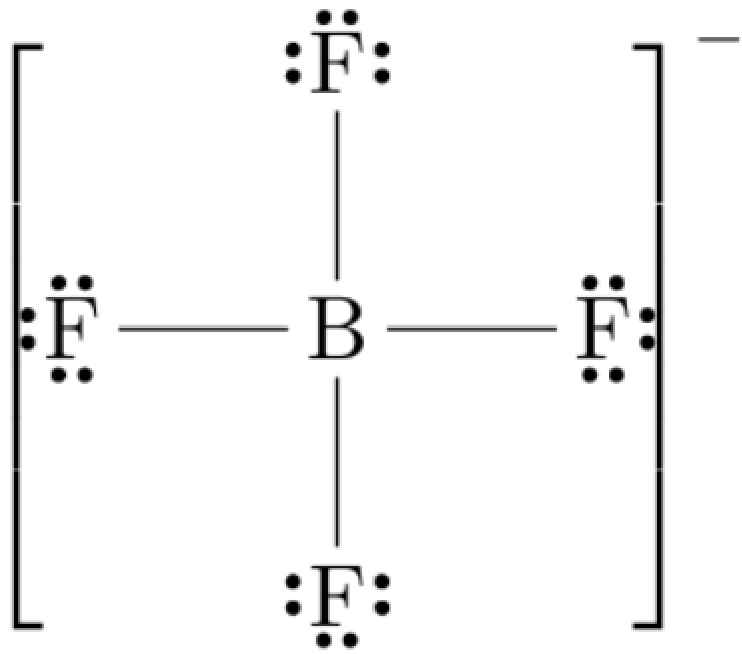
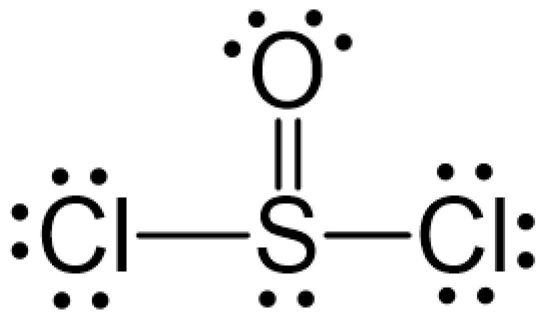
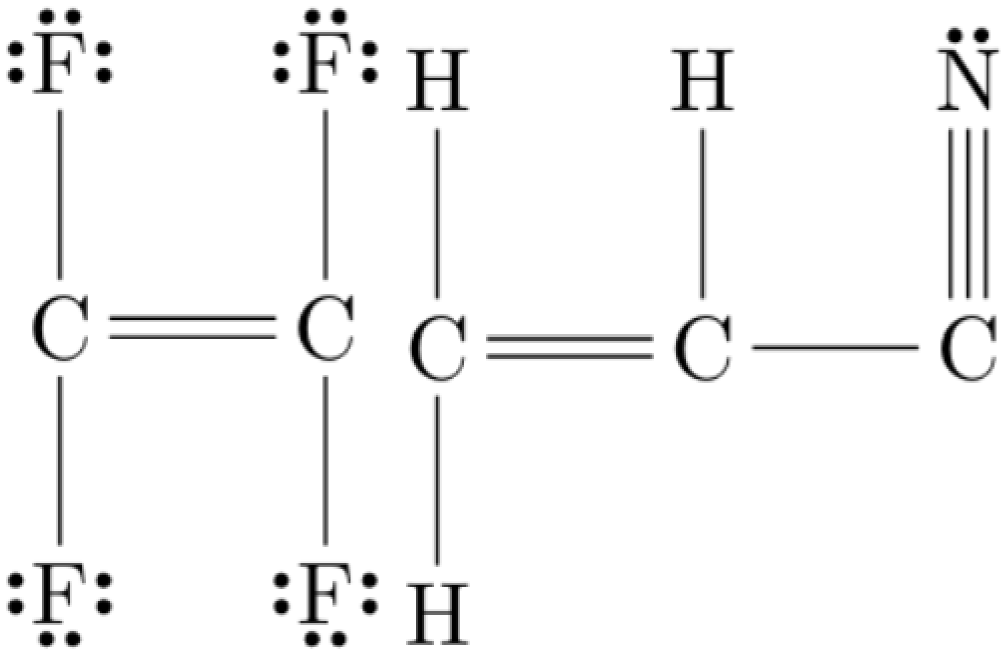






7.39 

7.40 It is impossible for hydrogen to be the central atom in the Lewis structure of a polyatomic molecule, because its outer shell can only have a maximum of two electrons. Therefore, hydrogen can only form one bond.