1. a.

b. There are eight:

\[
\begin{align*}
&\text{F}_3\text{C}\text{C}\text{F}_3 \quad \text{F}_3\text{C}\text{O}\text{F} \\
&\text{F}_3\text{C}\text{O} \quad \text{F}_3\text{C}\text{F}_3 \\
&\text{F}_3\text{C}\text{F}_3 \quad \text{F}_3\text{C}\text{O} \\
&\text{F}_3\text{C}\text{O} \quad \text{F}_3\text{C}\text{F}_3
\end{align*}
\]
2.

a) ![Lewis structure a]

b) ![Lewis structure b]

c) ![Lewis structure c]

Notice that these Lewis structures have been drawn as line-bond structures, and incorporate bond angles appropriate to the hybridization states of their atoms. It is typical for organic chemists to draw Lewis line-bond structures this way.

3. a)

\[
\begin{align*}
2p & \uparrow \uparrow \downarrow \downarrow \\
2s & \uparrow \uparrow \\
F^{1-} & 1s \uparrow \uparrow
\end{align*}
\]

b)

\[
\begin{align*}
2p & \downarrow \downarrow \downarrow \\
2s & \uparrow \uparrow \\
N & 1s \downarrow \uparrow
\end{align*}
\]

c)

\[
\begin{align*}
3d & \_ \_ \_ \_ \_ \_ \\
3p & \_ \_ \_ \\
3s & \_ \\
2p & \uparrow \uparrow \uparrow \uparrow \\
2s & \uparrow \uparrow \\
Al^{3+} & 1s \uparrow \uparrow
\end{align*}
\]

Organic chemists normally draw electronic configurations of atomic species using only the core and valence orbitals of atomic species: that’s why, for example, \( F^{1-} \) above was configured without 3s, 3p, 3d, or higher orbitals.
4. a) 

\[
\begin{align*}
3d & \quad \quad \quad \quad \quad \quad \quad \quad \\
3p & \quad \quad \quad \quad \\
3s & \quad \\
2p & \quad \quad \quad \quad \quad \quad \quad \\
2s & \quad \\
\text{Mg}^{2+} & 1s \quad \quad \quad \\
\end{align*}
\]

b) 

\[
\begin{align*}
2p & \quad \quad \quad \quad \\
2s & \quad \quad \quad \\
\text{C}^{+} & 1s \quad \quad \quad \\
\end{align*}
\]

c) 

\[
\begin{align*}
2sp^3 & \quad \quad \quad \quad \quad \quad \quad \\
\text{O}^{+} & 1s \quad \quad \quad \quad \quad \quad \\
\end{align*}
\]

5. a) 

\[
\begin{align*}
2sp^3 & \quad \quad \quad \quad \quad \quad \quad \quad \\
\text{B}^{-} & 1s \quad \quad \quad \quad \quad \quad \quad \quad \\
\end{align*}
\]

b) 

\[
\begin{align*}
2p & \quad \quad \quad \quad \\
2sp^2 & \quad \quad \quad \quad \quad \quad \quad \quad \\
\text{O}^{+} & 1s \quad \quad \quad \quad \quad \quad \quad \quad \\
\end{align*}
\]
b) Less stable: $\text{He}_2$ has two electrons in an antibonding orbital, whereas 2 He have none. The rule of thumb is that bonding interactions are better than nonbonding interactions, which in turn are better than antibonding interactions:

$$E_\sigma > E_n > E_{\sigma^*}$$
7. a) 

b) Five: \( \text{Csp}^3 - \text{H}1s, \ \text{Nsp}^3 - \text{H}1s, \ \text{Csp}^3 - \text{Nsp}^3, \ \text{Nsp}^3 - \text{Osp}^3, \ \text{Osp}^3 - \text{H}1s \)

c) \( < 109.5 \, ^\circ \)

8. a) 

b) Six: \( \sigma \text{Csp}^3 - \text{H}1s, \ \sigma \text{Nsp} - \text{Csp}, \ \pi \text{Nsp} - \text{Csp}, \ \sigma \text{Csp}^3 - \text{Csp}, \ \sigma \text{Csp}^3 - \text{Osp}^3, \ \sigma \text{Osp}^3 - \text{H}1s \)

c) \( < 109.5 \, ^\circ \)
9. a) 

b) Seven: 2 different C\textit{sp}^3—C\textit{sp}^3, 3 different C\textit{sp}^3—H1\textit{s}, and 2 different C\textit{sp}^3—O\textit{sp}^3.

10. a) Si(CH\textsubscript{3})\textsubscript{4} is a bigger “ball” (larger surface area for induced dipole-induced dipole effects) than C(CH\textsubscript{3})\textsubscript{4}. Therefore, it has the higher boiling point.

b) Si—C bonds are too long to allow for the good Si\textit{p}—C\textit{p} orbital overlap needed to create a \pi bond.

11. a) 

b) 14

c) about 109.5\degree
12. a. 

\[
\begin{array}{cccc}
\text{Na}^+ & \text{F}^+ & \text{Ca}^{2+} & \text{O}^{2-} \\
\text{Cl}^- & \text{H}^- & \text{S}^{2-} & \text{Br}^+ \\
\end{array}
\]

b. 

\[
\begin{array}{c}
2p \\
2sp \\
F^+ 1s
\end{array}
\]

13. 

hybridization state: \( sp^3 \) 

hybridization state: \( sp^2 \)

a. Clearly add lone pair electrons so as to fill the octets of all atoms of caffeine.

b. In the spaces provided, provide the hybridization state of the nitrogens indicated.

c. The bond that the squiggly arrow is pointed at is at a double bond. This bond has formed from overlap of a \( sp^2 \) orbital of carbon and a \( sp^2 \) orbital of nitrogen to form a \( \sigma \) bond, and from overlap of a \( p \) orbital of nitrogen and a \( p \) orbital of carbon to form a \( \pi \) bond.
14.

a. 
\[
\begin{align*}
&\text{CH}_3\text{C}=\text{H} > \text{CH}_3\text{C}^\text{\_} \text{S}^\text{\_} \text{H} > \text{CH}_3\text{C} = \text{H} \\
&\text{H}_3\text{C} = \text{H} > \text{CH}_3\text{C}^\text{\_} \text{S}^\text{\_} \text{H} > \text{CH}_3\text{C} = \text{H}
\end{align*}
\]

b. It’s the C–C bond.

c. It’s the O–O bond.

15.

A         B          C      D           E

> >
16. This is a combination of two effects. First, bond polarization is a factor of both electronegativity difference and the length of the bond: the longer the bond between two different atoms, the more polarized that bond is. A C–Cl bond is 30% longer than a C–F bond, so the effective polarizing ability of Cl on C is about one-third greater than it would be if the C–Cl bond length were the same as a C–F bond length. Put another way, for C–Cl bonds, the electronegativity of Cl that C “feels” really is one-third greater than atomic chlorine’s value (4.0 instead of 3.0). Chlorine has about the same polarizing effect on carbon as fluorine because of the difference in bond lengths!

Second, being in row 3 of the periodic table, chlorine is bigger than fluorine. It occupies a larger volume of space than fluorine, so it has a greater effect on compression of the bond angles \( \angle \text{HCH} \) in \( \text{CH}_3\text{Cl} \) compared to \( \text{CH}_3\text{F} \), effectively increasing the contribution of H–C bond polarization to the dipole moment of \( \text{CH}_3\text{Cl} \) relative to \( \text{CH}_3\text{F} \).
17.  

a.  

\[ 2p \quad _____ \]

\[ O \quad sp^2 \quad \uparrow \quad \uparrow \quad \uparrow \quad \uparrow \]

b.  a nonpolar covalent bond

c.  carbon or iodine