4. (E) For simplicity, the 3 lone pairs on halogens are not shown in the structures below. Add 6 electrons per halogen atom.

(a) $\text{O} = \text{C} = \text{S}$

(b) $\text{H} \quad \text{O} \quad \text{H}
\text{H} \quad \text{C} \quad \text{C} \quad \text{H}$

(c) $\text{F} \quad \text{C} = \text{O}$

(d) $\text{Cl} \quad \text{S} \quad \text{O}$

(e) $\text{H} \quad \text{C} = \text{C} \quad \text{H}$

11. (M) The answer is (c), hypochlorite ion. The flaws with the other answers are as follows:

(a) $\text{O} = \text{C} = \text{N}$ does not have an octet of electrons around C.

(b) $[\text{C} = \text{C}]^-$ does not have an octet around either C. Moreover, it has only 6 valence electrons in total while it should have 10, and finally, the sum of the formal charges on the two carbons doesn’t equal the charge on the ion.

(d) The total number of valence electrons in NO is incorrect. No, being an odd-electron species should have 11 valence electrons, not 12.

21. (E) $\text{FC} = \# \text{valence e}^- \text{ in free atom} - \# \text{number lone-pair e}^- - \frac{1}{2} \# \text{bond pair e}^- \text{e}^-$

(a) Central O in $\text{O}_3$: $6 - 2 - 3 = +1$

(b) Al in $\text{AlH}_4^-$: $3 - 0 - 4 = -1$

(c) Cl in $\text{ClO}_3$ : $7 - 2 - 5 = 0$

(d) Si in $\text{SiF}_6^{2-}$ : $4 - 0 - 6 = -2$

(e) Cl in $\text{ClF}_3$ : $7 - 4 - 3 = 0$

25. (E)

(a) $\text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H}$

(b) $\text{H} \quad \text{O} \quad \text{C} = \text{O}$

(c) $\text{O} \quad \text{H} \quad \text{O} \quad \text{S} \quad \text{S} \quad \text{OH}$
30. (E) In \( \text{C}_2\text{O}_2 \) there are \( 3 \times 4 + 2 \times 6 = 24 \) valence electrons or 12 valence electron pairs. A plausible Lewis structure follows: \( \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} \)

33. (E) (a) \( \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{H}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{H}}} \)

(b) \( \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{H}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{C}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}=\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{H}}} \)

37. (M) \( \text{Na}—\text{Cl} \) and \( \text{K}—\text{F} \) both possess bonds between a metal and a nonmetal. Thus, they have the largest ionic character, with the ionic character of \( \text{K}—\text{F} \) being greater than that of \( \text{Na}—\text{Cl} \), both because \( \text{K} \) is more metallic (closer to the lower left of the periodic table) than \( \text{Na} \) and because \( \text{F} \) is more nonmetallic (closer to the upper right) than \( \text{Cl} \). The remaining three bonds are covalent bonds to \( \text{H} \). Since \( \text{H} \) and \( \text{C} \) have about the same electronegativity (a fact you need to memorize), the \( \text{H}—\text{C} \) bond is the most covalent (or the least ionic). \( \text{Br} \) is somewhat more electronegative than \( \text{C} \), while \( \text{F} \) is considerably more electronegative than \( \text{C} \), making the \( \text{F}—\text{H} \) bond the most polar of the three covalent bonds. Thus, ranked in order of increasing ionic character, these five bonds are:

\[ \text{C}—\text{H} < \text{Br}—\text{H} < \text{F}—\text{H} < \text{Na}—\text{Cl} < \text{K}—\text{F} \]

The actual electronegativity differences follow:

\[ \Delta \text{EN} = 0.4 \quad 0.7 \quad 1.9 \quad 2.1 \quad 3.2 \]

48. (M) The only one of the four species that requires resonance forms to correctly describe the intramolecular bonding is \( \text{CO}_3^{2-} \). Resonance forms of equal energy cannot be generated for the other species. All four Lewis structures are drawn below.
(a) In CO$_2$, there are 4+2X6=16 valence electrons, or 8 electron pairs. \( \overset{\cdot}{O} = C = \overset{\cdot}{O} \)

(b) In OCl$^-$, there are 6 + 7 + 1 = 14 valence electrons, or 7 electron pairs, \( \overset{\cdot}{O} = \overset{\cdot}{C} - \overset{\cdot}{Cl} \).

(c) In CO$_3^{2-}$, there are 4+3X6=24 valence electrons, or 12 electron pairs (see below).

\[
\begin{align*}
\overset{\cdot}{O} = C - \overset{\cdot}{O} \quad &\leftrightarrow\quad \overset{\cdot}{O} = C - \overset{\cdot}{O} \\
\end{align*}
\]

(d) In OH$^-$, there are 6 + 1 + 1 = 8 valence electrons, or 4 electron pairs. \( \overset{\cdot}{O} - \overset{\cdot}{H} \)

53. (M)

(a) CH$_3$ has a total of 3X1+4=7 valence electrons, or 3 electron pairs and a lone electron. C is the central atom. A plausible Lewis structure is shown on the right.

(b) ClO$_2$ has a total of 2X6+7=19 valence electrons, or 9 electron pairs and a lone electron. Cl is the central atom. A plausible Lewis structure is: \( \overset{\cdot}{O} - \overset{\cdot}{Cl} - \overset{\cdot}{O} \).

(c) NO$_3$ has a total of 3X6+5=23 valence electrons, or 11 electron pairs, plus a lone electron. N is the central atom. A plausible Lewis structure is shown to the right. Other resonance forms can also be drawn. \( \overset{\cdot}{O} = N - \overset{\cdot}{O} \).

57. (M) In PO$_4^{3-}$: 5+4X6+3=32 valence electrons or 16 electron pairs. An expanded octet is not needed.

In PI$_3$: 5+3X7=26 valence electrons or 13 electron pairs. An expanded octet is not needed.

In ICl$_3$: 7+3X7=28 valence electrons or 14 electron pairs. An expanded octet is necessary.

In OSCl$_2$: 6+6+2X7=26 valence electrons or 13 electron pairs. An expanded octet is not needed.

In SF$_4$: 6+4X7=34 valence electrons or 17 electron pairs. An expanded octet is necessary.

In ClO$_4^-$: 7+4X6+1=32 valence electrons or 16 electron pairs. An expanded octet is not needed.