1) Let’s start with a simple MO diagram for a second row diatomic:

```
  ___  
 /    
\    
 2p   2p
 /    \\
\    
 2s   2s
```

Carbon atoms have the configuration: C: [He]2s²2p² so \( C_{2}^{-} \) would have 10 electrons (need the extra two for the 2 negative charges). We have a configuration which looks like: 

\[
(\sigma_{s})^{2} (\sigma_{s}^{*})^{2} (\pi)^{4} (\sigma_{p})^{2},
\]

which leads to a bond order of: 

\[
\frac{1}{2} (8-2) = 3.
\]

To describe the bonding, one might say that there are two \( \pi \) bonds, and one \( \sigma \) bond, since the \( \sigma_{p} \) and both \( \pi_{p} \) orbitals are fully occupied. Given this last fact, all electrons are paired, and the anion would be predicted to be diamagnetic.

2) We can use a diagram similar to the one below for both \( O_{2} \) and \( N_{2} \) molecules:

```
  ___  
 /    
\    
 2p   2p
 /    \\
\    
 2s   2s
```

Recall that the ground-state electron configurations of \( N \) and \( O \) are: \( N: [He] 2s^{2} 2p^{3} \) and \( O: [He] 2s^{2} 2p^{4} \). In \( N_{2} \), the orbital occupancies are: 

\[
(\sigma_{s})^{3} (\sigma_{s}^{*})^{3} (\pi)^{3} (\sigma_{p})^{2}
\]

whereas in \( O_{2} \), the orbital occupancies are: 

\[
(\sigma_{s})^{3} (\sigma_{s}^{*})^{3} (\pi)^{3} (\sigma_{p})^{2} (\pi^{*})^{2}.
\]

The ionization
energy would be the energy needed to remove an electron from the highest occupied orbital. In N\(_2\) the highest occupied (molecular) orbital is stabilized relative to the atomic orbitals from which it is composed, so it will take more energy to ionize molecular nitrogen (N\(_2\)) than it would to ionize atomic nitrogen (N). On the other hand, the highest occupied (molecular) orbital in O\(_2\) is destabilized relative to the atomic orbital of which it is composed; therefore, it will take less energy to remove an electron from molecular oxygen (O\(_2\)) than from atom oxygen (O).

3) We don’t have much to work with, but let’s go with what we do have. We are told that 423 mL of Br\(_2\)(g) at 150°C and 748 mmHg (= 0.984 atm) produces 3.28 g BrF\(_n\); we have, therefore, from PV=nRT, 

\[
n = \frac{\left(\frac{748 \text{ mmHg}}{760 \text{ mmHg/atm}}\right)(0.423 L)}{(0.08206 L \cdot \text{atm/K/mol})(423 K)} = 1.20 \times 10^{-2} \text{ mol Br}_2.
\]

We can now write the following ‘balanced’ equation:

\[
\text{Br}_2 + nF_2 \rightarrow 2\text{BrF}_n
\]

which tells us that every one mole of Br\(_2\) produces two moles of BrF\(_n\), so we know that 4.20 g BrF\(_n\) represents 2.40 x 10\(^{-2}\) moles, and therefore, the M\(_m\) = \(\frac{3.28g}{2.40 \times 10^{-2} \text{ mol}} = 136.74 \text{ g/mol}\). The atomic mass of Br is 79.904, and that of F is 18.998. We know that of the 175 g/mol. molecular mass, 79.904 g is taken up with the Br, so the rest (136.7-79.9) = 56.8 g is the F. The number of fluorines, \(n\), in the molecule is therefore: 

\[
n = \frac{56.8g}{18.998g/\text{mol}} = 3,
\]

and the formula is: BrF\(_3\). The geometry is:

![BrF3 structure](image)

(that is, T-shaped for the molecule; trigonal bipyramidal for the electron pairs), which leads to hybrid orbitals, under the VBT approach, of dsp\(^3\).

4a) Here are our two options for N\(_2\)O:
The second option is not polar, therefore, the first structure is the correct one.

b) The two good resonance forms are:

For the structure on the left: the left most nitrogen has a formal charge of \(-1\); the middle nitrogen, as shown, is +1; the oxygen is, as shown, 0. For the structure on the right, the left-most nitrogen has a formal charge of 0; the middle nitrogen is (again) +1; and the oxygen has a formal charge of -1. (You may have found some other, less good resonance forms, and I will look at these too.) The central nitrogen in both has two electron domains around it (from the VSEPR model), and therefore, being linear, should exhibit \(sp + p + p\) hybridization

c) This is the right-most resonance form above, and so would be described in this manner:

The left-most nitrogen has two pairs around it, and is therefore \(sp + p + p\) hybridized; the central nitrogen also has two pairs around it and is also \(sp + p + p\) hybridized; the oxygen has four pairs around it, and should exhibit \(sp^3\) hybridization. That is, the nitrogens are \(\sigma\)-bonded using sp hybrids, and make two \(\pi\)-bonds using their p-orbitals. The lone pair on the left nitrogen is in a sp hybrid orbital. The oxygen bonds to the central nitrogen via an sp-sp\(^3\) \(\sigma\)-bond.

5) We are looking at the following reaction:

\[
\text{CH}_3\text{CH}_3 \rightarrow 2 \text{CH}_3^* 
\]

that is, a methyl radical. It is an odd-electron species, having only 7 electrons in its valence shell. None of our rules have explicitly dealt with situations like this one, so we have to make some reasonable guesses and extrapolations. We know two things:

1) \(\text{CH}_3^+\) (the cation) would have only 6 electrons in the valence shell, composed of three bonding pairs, and no lone pairs. VSEPR would predict, therefore, a trigonal planar geometry, for which the bond angle would be 120°.
2) CH₃⁻ (the anion) has 8 electrons in the valence shell, composed of three bonding pairs, and one lone pair. VSEPR would predict, therefore, an ideal tetrahedral geometry, for which the bond angle would be 109.47° (or actually, somewhat less, since the lone pair would occupy more room.)

The cation has an empty p-orbital on carbon, and the anion has a filled sp³ orbital (under the VBT model) on carbon. The radical has a half-filled orbital on carbon, and would therefore be expected to be intermediate between the anion and the cation; this said, we might comfortably predict a bond angle of 115° or so. (Mostly anything between 109.5 and 120 (not inclusive) is acceptable, given that you support your answer.) (Experimental data on this species lead to a prediction of 117°, so 115° was not too far off.)