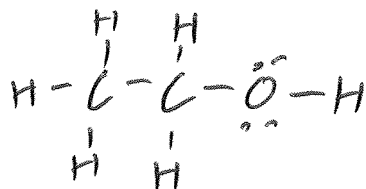
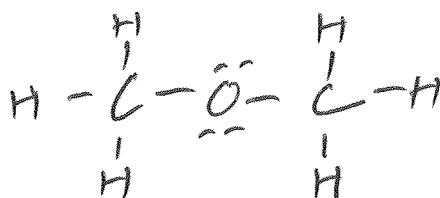


06/18/24

Draw the Lewis structure for ethyl alcohol which has a molecular formula of C_2H_6O and a condensed formula of CH_3CH_2OH .



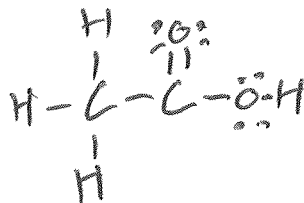
Draw the Lewis structure for acetone which has a molecular formula of C_3H_6O and a condensed formula of CH_3COCH_3 .



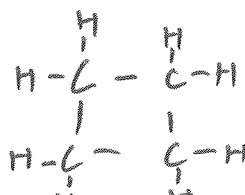
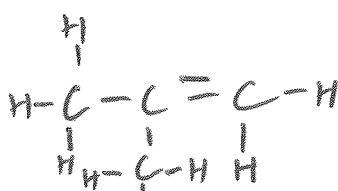
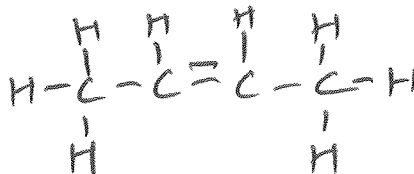
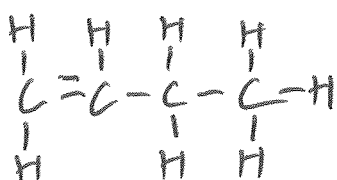
Draw the Lewis structure for carbon dioxide (CO_2)



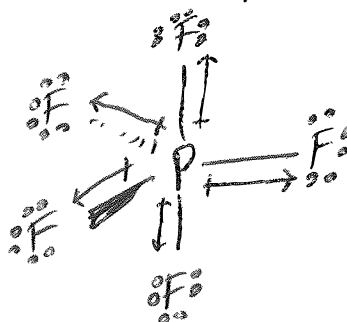
Draw the Lewis structure for acetic acid which has a molecular formula of $C_2H_4O_2$ and a condensed formula of CH_3COOH .



Draw 4 different Lewis structures that all have the molecular formula of C_4H_8

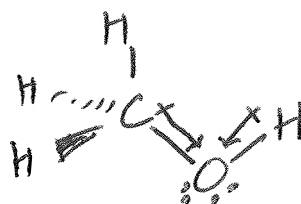
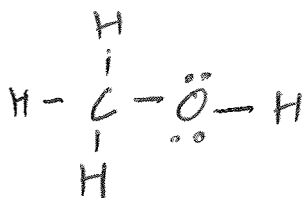


Draw the Lewis structure and then the 3-dimensional structure for the PF_5 molecule. Indicate any polar bonds and indicate whether the whole molecule is polar.



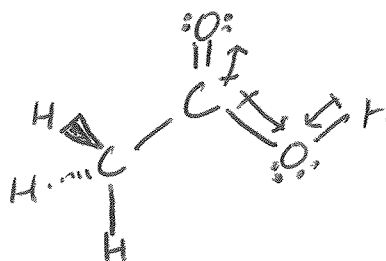
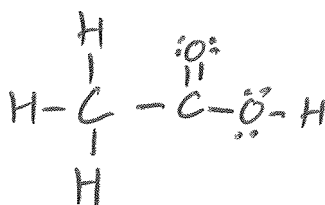
molecule is
non-polar

Draw the Lewis structure and then the 3-dimensional structure for the methanol (CH_3OH) molecule. Indicate any polar bonds and indicate whether the whole molecule is polar.



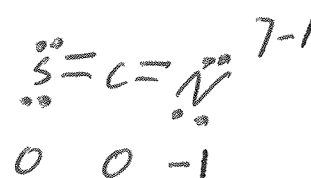
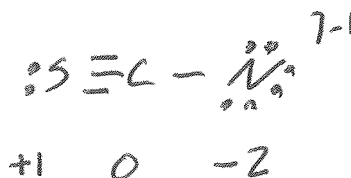
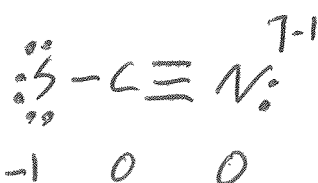
molecule is polar

Draw the Lewis structure and then the 3-dimensional structure for the acetic acid (CH_3COOH) molecule. Indicate any polar bonds and indicate whether the whole molecule is polar.



molecule is polar

Draw the possible resonance structures for the thiocyanate ion (SCN^{-1}) and use formal charges to decide which one is the best representation of the structure.



Best b/c fewest non-zero FC and -1 FC on more electroneg atom

Use the table of bond dissociation energies to determine the amount of energy released when 100.0 grams of ethyl alcohol ($\text{CH}_3\text{CH}_2\text{OH}$) is burned. Compare this to the amount of energy released when 100.0g of octane (C_8H_{18}) is burned.



<u>Break</u>	<u>Make</u>	$\left[(4 \times -745 \frac{\text{kJ}}{\text{mol}}) + (6 \times -463 \frac{\text{kJ}}{\text{mol}}) \right] = -5758 \frac{\text{kJ}}{\text{mol rxn}} \text{ Making}$ $\left[(5 \times 413 \frac{\text{kJ}}{\text{mol}}) + (3 \times 498 \frac{\text{kJ}}{\text{mol}}) + (1 \times 346 \frac{\text{kJ}}{\text{mol}}) + (1 \times 358 \frac{\text{kJ}}{\text{mol}}) + (1 \times 463 \frac{\text{kJ}}{\text{mol}}) \right]$ $= 4726 \frac{\text{kJ}}{\text{mol rxn}} \text{ Breaking}$ $\Delta H = -1032 \frac{\text{kJ}}{\text{mol rxn}}$ $100.0 \text{g C}_2\text{H}_6\text{O} \times \frac{1 \text{mol}}{46 \text{g}} \times \frac{-1032 \text{kJ}}{\text{mol}} = -2243 \text{ kJ}$
5 C-H	4 C=O	
3 O=O	6 O-H	
1 C-C		
1 C-O		
1 O-H		



<u>Break</u>	<u>Make</u>	$\left[(16 \times -745 \frac{\text{kJ}}{\text{mol}}) + (18 \times -463 \frac{\text{kJ}}{\text{mol}}) \right] = -20254 \frac{\text{kJ}}{\text{mol rxn}} \text{ Making}$ $\left[(18 \times 413 \frac{\text{kJ}}{\text{mol}}) + (17 \times 346 \frac{\text{kJ}}{\text{mol}}) + (12.5 \times 498 \frac{\text{kJ}}{\text{mol}}) \right] = 16081 \frac{\text{kJ}}{\text{mol rxn}} \text{ Breaking}$ $\Delta H = -4173 \text{ kJ/mol rxn}$ $100.0 \text{g C}_8\text{H}_{18} \times \frac{1 \text{mol}}{114 \text{g}} \times \frac{-4173 \text{kJ}}{\text{mol}} = -3660 \text{ kJ}$
18 C-H	16 C=O	
7 C-C	18 O-H	
$\frac{25}{2} \text{O}_2$		

Ethyl alcohol is used as a fuel additive in gasoline (which is pretty much just octane). The density of ethyl alcohol is 0.79 g/cm^3 and the density of octane is 0.70 g/cm^3 . Is an ethanol/gasoline mixture a better fuel than just pure gasoline? Explain.

$$\text{Ethanol} \quad 0.79 \frac{\text{g}}{\text{cm}^3} \times \frac{-2243 \text{ kJ}}{100 \text{ g}} = -17.72 \text{ kJ/cm}^3$$

$$\text{Octane} \quad 0.70 \frac{\text{g}}{\text{cm}^3} \times \frac{-3660 \text{ kJ}}{100 \text{ g}} = -25.62 \text{ kJ/cm}^3$$

Replacing octane w/ equal volume of ethanol
decreases energy content of fuel