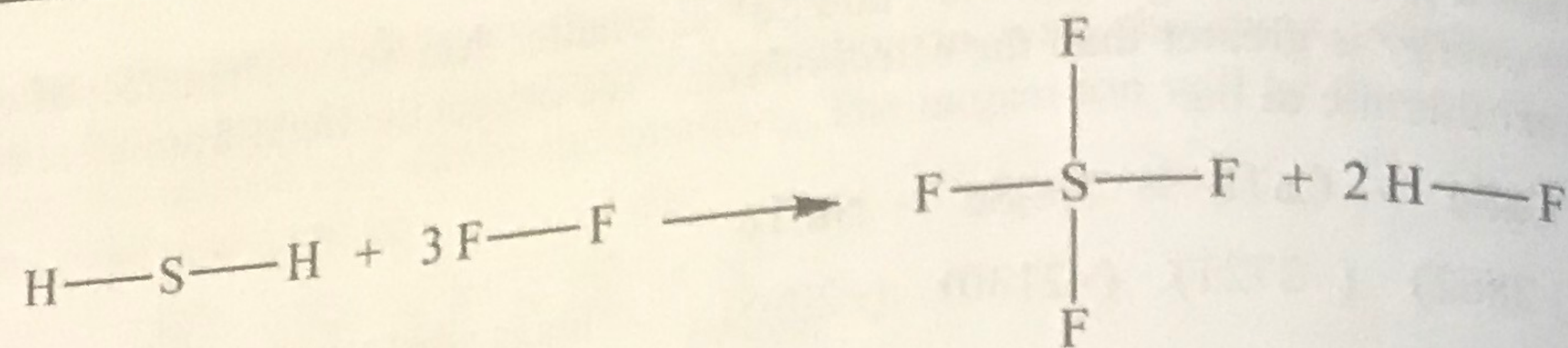


67.



Bonds broken:

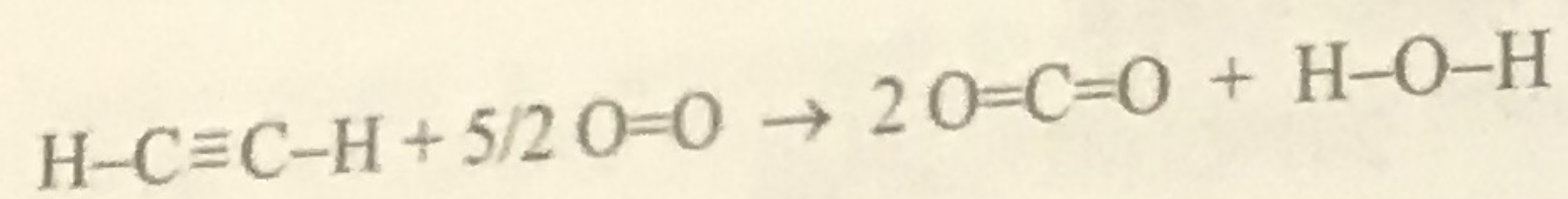
2 S-H (347 kJ/mol)
3 F-F (154 kJ/mol)

Bonds formed:

4 S-F (327 kJ/mol)
2 H-F (565 kJ/mol)

$$\Delta H = 2(347) + 3(154) - [4(327) + 2(565)] = -1282 \text{ kJ}$$

69.



Bonds broken:

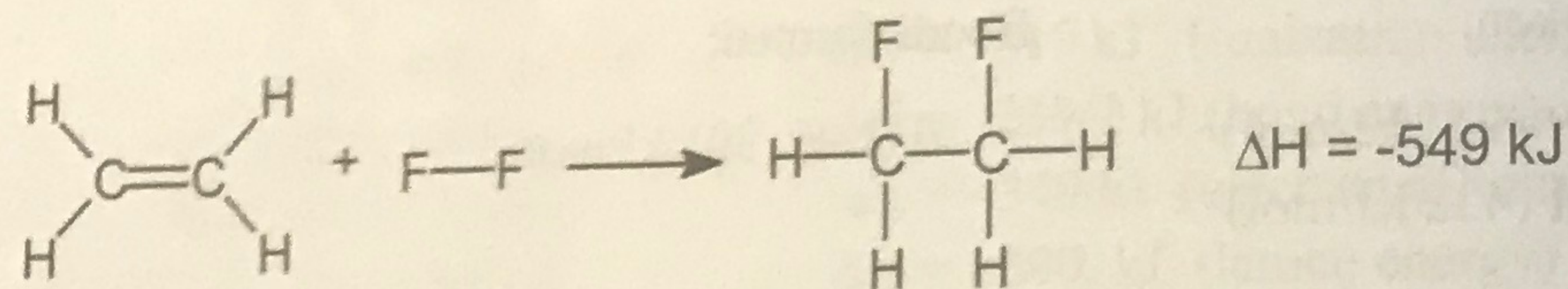
2 C-H (413 kJ/mol)
1 C≡C (839 kJ/mol)
5/2 O=O (495 kJ/mol)

Bonds formed:

2 × 2 C=O (799 kJ/mol)
2 O-H (467 kJ/mol)

$$\Delta H = 2(413 \text{ kJ}) + 839 \text{ kJ} + 5/2 (495 \text{ kJ}) - [4(799 \text{ kJ}) + 2(467 \text{ kJ})] = -1228 \text{ kJ}$$

71.



Bonds broken:

1 C=C (614 kJ/mol)
1 F-F (154 kJ/mol)

Bonds formed:

1 C-C (347 kJ/mol)
2 C-F ($D_{\text{CF}} = \text{C-F bond energy}$)

$$\Delta H = -549 \text{ kJ} = 614 \text{ kJ} + 154 \text{ kJ} - [347 \text{ kJ} + 2D_{\text{CF}}], \quad 2D_{\text{CF}} = 970., \quad D_{\text{CF}} = 485 \text{ kJ/mol}$$

73. a. $\Delta H^\circ = 2\Delta H_{\text{f,HCl}}^\circ = 2 \text{ mol}(-92 \text{ kJ/mol}) = -184 \text{ kJ}$ (= -183 kJ from bond energies)

b. $\Delta H^\circ = 2\Delta H_{\text{f,NH}_3}^\circ = 2 \text{ mol}(-46 \text{ kJ/mol}) = -92 \text{ kJ}$ (= -109 kJ from bond energies)

Comparing the values for each reaction, bond energies seem to give a reasonably good estimate for the enthalpy change of a reaction. The estimate is especially good for gas phase reactions.

75. a. Using SF₄ data: $\text{SF}_4(\text{g}) \rightarrow \text{S}(\text{g}) + 4 \text{F}(\text{g})$

$$\Delta H^\circ = 4D_{\text{SF}} = 278.8 + 4(79.0) - (-775) = 1370. \text{ kJ}$$

$$D_{\text{SF}} = \frac{1370. \text{ kJ}}{4 \text{ mol SF bonds}} = 342.5 \text{ kJ/mol} = \text{S-F bond energy}$$

Using SF₆ data: $\text{SF}_6(\text{g}) \rightarrow \text{S}(\text{g}) + 6 \text{F}(\text{g})$

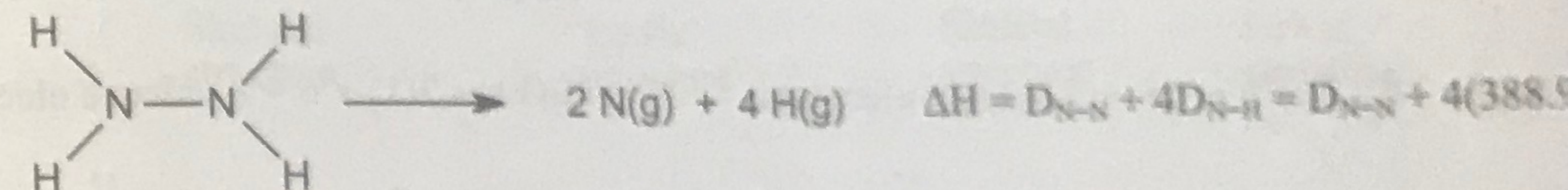
$$\Delta H^\circ = 6D_{\text{SF}} = 278.8 + 6(79.0) - (-1209) = 1962 \text{ kJ}$$

$$D_{\text{SF}} = \frac{1962. \text{ kJ}}{6 \text{ mol}} = 327.0 \text{ kJ/mol} = \text{S-F bond energy}$$

b. The S-F bond energy in the table is 327 kJ/mol. The value in the table was based on the S-F bond in SF₆.

c. S(g) and F(g) are not the most stable forms of the elements at 25°C. The most stable forms are S₈(s) and F₂(g); $\Delta H_{\text{f}}^\circ = 0$ for these two species.

77.



$$\Delta H^\circ = 2\Delta H_{\text{f,N}}^\circ + 4\Delta H_{\text{f,H}}^\circ - \Delta H_{\text{f,N}_2\text{H}_4}^\circ = 2(472.7 \text{ kJ}) + 4(216.0 \text{ kJ}) - 95.4 \text{ kJ}$$

$$\Delta H^\circ = 1714.0 \text{ kJ} = D_{\text{N-N}} + 4(388.9)$$

$$D_{\text{N-N}} = 158.4 \text{ kJ/mol} \text{ (versus } 160. \text{ kJ/mol in Table 8.4)}$$

Lewis Structures and Resonance

79. Drawing Lewis structures is mostly trial and error. However, the first two steps are the same. These steps are (1) count the valence electrons available in the molecule/ion the same. These steps are (1) count the valence electrons available in the molecule/ion the same. These steps are (1) count the valence electrons available in the molecule/ion the same. (2) attach all atoms to each other with single bonds (called the skeletal structure) noted otherwise, the atom listed first is assumed to be the atom in the middle, called central atom, and all other atoms in the formula are attached to this atom. The most exceptions to the rule are formulas that begin with H, e.g., H₂O, H₂CO, etc. Hydrogen never be a central atom since this would require H to have more than two electrons. compounds, the atom listed second is assumed to be the central atom.

After counting valence electrons and drawing the skeletal structure, the rest is trial and error. We place the remaining electrons around the various atoms in an attempt to satisfy the octet rule (or duet rule for H). Keep in mind that practice makes perfect. After practicing, (and will) become very adept at drawing Lewis structures.