Section 5.3: Bond Energies

Tutorial 1 Practice, page 312

1. There are 3 C–H bonds and 1 C–Cl bond in chloromethane, $CH_3Cl(g)$.

Given: $n_{\text{C–H}} = 3$ mol; $D_{\text{C–H}} = 413$ kJ/mol; $n_{\text{C–Cl}} = 1$ mol; $D_{\text{C–Cl}} = 339$ kJ/mol

Required: ΔH

Analysis: $\Delta H = \sum n \times D_{\text{bonds broken}}$

Solution:

$$\begin{aligned}\Delta H &= \sum n \times D_{\text{bonds broken}} \\ &= 3\ \text{mol} \times D_{\text{C–H}} + 1\ \text{mol} \times D_{\text{C–Cl}} \\ &= 3\ \cancel{\text{mol}} \times \frac{413\ \text{kJ}}{\cancel{\text{mol}}} + 1\ \cancel{\text{mol}} \times \frac{339\ \text{kJ}}{\cancel{\text{mol}}}\end{aligned}$$

$\Delta H = 1578$ kJ

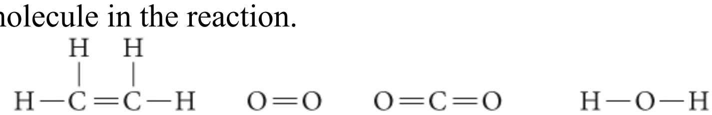
Statement: The energy required to separate 1 mol of chloromethane into free atoms is 1578 kJ.

2. Solution:

Step 1: Write the balanced chemical equation for the combustion of 1 mol of ethene gas.

$C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(l)$

Step 2: Determine the bonding of each substance by drawing structural formulas for each molecule in the reaction.



ethene oxygen carbon dioxide water

Step 3: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	**Number of bonds per mole ($n_{\text{substance}}$)**	**Amount of bonds in reaction**	**Bond energy per mole**
reactants	C_2H_4	4 mol C–H bonds 1 mol C=C bonds	4 mol 1 mol	413 kJ/mol 614 kJ/mol
	O_2	1 mol O=O bonds	3 mol	495 kJ/mol
products	CO_2	2 mol C=O bonds	4 mol	799 kJ/mol
	H_2O	2 mol O–H bonds	4 mol	467 kJ/mol

Step 4: Calculate the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 4\text{ mol} \times D_{\text{C-H}} + 1\text{ mol} \times D_{\text{C=C}} + 3\text{ mol} \times D_{\text{O=O}} -$$

$$\left(4\text{ mol} \times D_{\text{C=O}} + 4\text{ mol} \times D_{\text{O-H}}\right)$$

$$= \left(4\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{614\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(3\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right) -$$

$$\left[\left(4\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(4\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 1652\text{ kJ} + 614\text{ kJ} + 1485\text{ kJ} - (3196\text{ kJ} + 1868\text{ kJ})$$

$$\Delta H = -1313\text{ kJ}$$

Statement: Since enthalpy change, ΔH, of the reaction is negative, the complete combustion of ethene gas to gaseous carbon dioxide and liquid water is an exothermic reaction.

3. Solution:

Step 1: Use the balanced chemical equation:

$N_2H_2(g) + F_2(g) \rightarrow N_2(g) + 2\ HF(g)$

Determine the bonding of each substance by drawing structural formulas for each molecule in the reaction.

$H-N=N-H$	$F-F$	$N \equiv N$	$H-F$
$N_2H_2(g)$	$F_2(g)$	$N_2(g)$	$HF(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	**Number of bonds per mole ($n_{\text{substance}}$)**	**Amount of bonds in reaction**	**Bond energy per mole**
reactants	N_2H_2	2 mol N–H bonds 1 mol N=N bonds	2 mol 1 mol	391 kJ/mol 418 kJ/mol
	F_2	1 mol F–F bonds	1 mol	154 kJ/mol
products	N_2	1 mol N≡N bonds	1 mol	941 kJ/mol
	HF	1 mol H–F bonds	2 mol	565 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of the reaction.

$$\begin{aligned}
\Delta H &= \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}} \\
&= 2\text{ mol} \times D_{\text{N-H}} + 1\text{ mol} \times D_{\text{N=N}} + 1\text{ mol} \times D_{\text{F-F}} - \\
&\quad \left(1\text{ mol } \times D_{\text{N}\equiv\text{N}} + 2\text{ mol} \times D_{\text{H-F}}\right) \\
&= \left(2\ \cancel{\text{mol}} \times \frac{391\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{418\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{154\text{ kJ}}{\cancel{\text{mol}}}\right) - \\
&\quad \left[\left(1\ \cancel{\text{mol}} \times \frac{941\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{565\text{ kJ}}{\cancel{\text{mol}}}\right)\right] \\
&= 782\text{ kJ} + 418\text{ kJ} + 154\text{ kJ} - (941\text{ kJ} + 1130\text{ kJ}) \\
\Delta H &= -717\text{ kJ}
\end{aligned}$$

Statement: When enthalpy change, ΔH, has a negative value, there is more energy released by the formation of the bonds in the products than is absorbed by the breaking of the bonds in the reactants. Therefore, thermal energy is released to the surroundings, and the reaction is exothermic.

4. Solution:

Step 1: Write the equation that represents the complete combustion of propane.

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	C_3H_8	8 mol C–H bonds	8 mol	413 kJ/mol
		2 mol C–C bonds	2 mol	347 kJ/mol
	O_2	1 mol O=O bonds	5 mol	495 kJ/mol
products	CO_2	2 mol C=O bonds	6 mol	799 kJ/mol
	H_2O	2 mol O–H bonds	8 mol	467 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of the reaction.

$$\begin{aligned}
\Delta H &= \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}} \\
&= 8\text{ mol} \times D_{\text{C-H}} + 2\text{ mol} \times D_{\text{C-C}} + 5\text{ mol} \times D_{\text{O=O}} - \\
&\quad \left(6\text{ mol} \times D_{\text{C=O}} + 8\text{ mol} \times D_{\text{H-O}}\right) \\
&= \left(8\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{347\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(5\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right) - \\
&\quad \left[\left(6\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(8\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right)\right] \\
&= 3304\text{ kJ} + 694\text{ kJ} + 2475\text{ kJ} - (4794\text{ kJ} + 3736\text{ kJ}) \\
\Delta H &= -2057\text{ kJ}
\end{aligned}$$

Statement: The quantity of energy released for the complete combustion of 1 mol of propane is 2057 kJ.

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1. (a) Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{substance}$)	Amount of bonds in reaction	Bond energy per mole
reactants	H_2	1 mol H–H bonds	1 mol	432 kJ/mol
	Cl_2	1 mol Cl–Cl bonds	1 mol	239 kJ/mol
products	HCl	1 mol H–Cl bonds	2 mol	427 kJ/mol

Step 3: Calculate the enthalpy change, Δ*H*, of the reaction.

$$\begin{aligned}\Delta H &= \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}} \\ &= 1\ \text{mol} \times D_{\text{H-H}} + 1\ \text{mol} \times D_{\text{Cl-Cl}} - 2\ \text{mol} \times D_{\text{H-Cl}} \\ &= \left(1\ \cancel{\text{mol}} \times \frac{432\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{239\ \text{kJ}}{\cancel{\text{mol}}}\right) - \left(2\ \cancel{\text{mol}} \times \frac{427\ \text{kJ}}{\cancel{\text{mol}}}\right) \\ &= 432\ \text{kJ} + 239\ \text{kJ} - 854\ \text{kJ} \\ \Delta H &= -183\ \text{kJ}\end{aligned}$$

Statement: The Δ*H* for the reaction is –183 kJ.

(b) Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$N_2(g) + 3H_2(g) \rightarrow 2\ NH_3(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{substance}$)	Amount of bonds in reaction	Bond energy per mole
reactants	N_2	1 mol N≡N bonds	1 mol	941 kJ/mol
	H_2	1 mol H–H bonds	3 mol	432 kJ/mol
products	NH_3	3 mol N–H bonds	6 mol	391 kJ/mol

Step 3: Calculate the enthalpy change, Δ*H*, of the reaction.

$$\begin{aligned}\Delta H &= \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}} \\ &= 1\ \text{mol} \times D_{\text{N}\equiv\text{N}} + 1\ \text{mol} \times D_{\text{H-H}} - 2\ \text{mol} \times D_{\text{N-H}} \\ &= \left(1\ \cancel{\text{mol}} \times \frac{941\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(3\ \cancel{\text{mol}} \times \frac{432\ \text{kJ}}{\cancel{\text{mol}}}\right) - \left(6\ \cancel{\text{mol}} \times \frac{391\ \text{kJ}}{\cancel{\text{mol}}}\right) \\ &= 941\ \text{kJ} + 1296\ \text{kJ} - 2346\ \text{kJ} \\ \Delta H &= -109\ \text{kJ}\end{aligned}$$

Statement: The Δ*H* for the reaction is –109 kJ.

2. (a) Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$HCN(g) + 2\ H_2(g) \rightarrow CH_3NH_2(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{substance}$)	Amount of bonds in reaction	Bond energy per mole
reactants	HCN	1 mol C–H bonds 1 mol C≡N bonds	1 mol 1 mol	413 kJ/mol 891 kJ/mol
	H_2	1 mol H–H bonds	2 mol	432 kJ/mol
products	CH_3NH_2	1 mol C–H bonds 2 mol C–N bonds 2 mol N–H bonds	3 mol 1 mol 2 mol	413 kJ/mol 305 kJ/mol 391 kJ/mol

Step 3: Calculate the enthalpy change, Δ*H*, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$
$$= 1 \text{ mol} \times D_{\text{C-H}} + 1 \text{ mol} \times D_{\text{C}\equiv\text{N}} + 2 \text{ mol} \times D_{\text{H-H}} -$$
$$\left(3 \text{ mol} \times D_{\text{C-H}} + 1 \text{ mol} \times D_{\text{C-N}} + 2 \text{ mol} \times D_{\text{N-H}}\right)$$
$$= \left(1 \cancel{\text{mol}} \times \frac{413 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1 \cancel{\text{mol}} \times \frac{891 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2 \cancel{\text{mol}} \times \frac{432 \text{ kJ}}{\cancel{\text{mol}}}\right) -$$
$$\left[\left(3 \cancel{\text{mol}} \times \frac{413 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1 \cancel{\text{mol}} \times \frac{305 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2 \cancel{\text{mol}} \times \frac{391 \text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$
$$= 413 \text{ kJ} + 891 \text{ kJ} + 864 \text{ kJ} - (1239 \text{ kJ} + 305 \text{ kJ} + 782 \text{ kJ})$$
$$\Delta H = -158 \text{ kJ}$$

Statement: The Δ*H* for the reaction is –158 kJ.

(b) Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$N_2H_4(g) + 2\ F_2(g) \rightarrow N_2(g) + 4\ HF(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{substance}$)	Amount of bonds in reaction	Bond energy per mole
reactants	N_2H_4	4 mol N–H bonds 1 mol N–N bonds	4 mol 1 mol	391 kJ/mol 160 kJ/mol
	F_2	2 mol F–F bonds	2 mol	154 kJ/mol
products	N_2	1 mol N≡N bonds	1 mol	941 kJ/mol
	HF	1 mol H–F bonds	4 mol	565 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 4 \text{ mol} \times D_{\text{N-H}} + 1 \text{ mol} \times D_{\text{N-N}} + 2 \text{ mol} \times D_{\text{F-F}} - (1 \text{ mol} \times D_{\text{N}\equiv\text{N}} + 4 \text{ mol} \times D_{\text{H-F}})$$

$$= \left(4 \cancel{\text{mol}} \times \frac{391 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1 \cancel{\text{mol}} \times \frac{160 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2 \cancel{\text{mol}} \times \frac{154 \text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(1 \cancel{\text{mol}} \times \frac{941 \text{ kJ}}{\cancel{\text{mol}}}\right) + \left(4 \cancel{\text{mol}} \times \frac{565 \text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 1564 \text{ kJ} + 160 \text{ kJ} + 308 \text{ kJ} - (941 \text{ kJ} + 2260 \text{ kJ})$$

$$\Delta H = -1169 \text{ kJ}$$

Statement: The ΔH for the reaction is –1169 kJ.

3. The ΔH values of chemical reactions found using bond energies are not always equal to values found by experiment because bond energies are reported as average bond energies. The bond dissociation energy of a given bond depends on the types of atoms and bonds in the same molecule. Since the energy of a given bond is affected by the number of atoms and bonds around it, slightly different molecules of the same composition will in reality have different bond energies and enthalpies.

4. Given: $\Delta H = -549 \text{ kJ}$; $D_{\text{B-B}} = 432 \text{ kJ/mol}$; $D_{\text{A-A}} = \frac{1}{2} D_{\text{A-B}}$

Required: bond energy of A_2, $D_{\text{A-A}}$

Analysis: $\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$

Solution:

Step 1: Use the equation to determine the bonding of each substance.

$A_2 + B_2 \rightarrow 2\ AB$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	A_2	1 mol A–A bonds	1 mol	$D_{\text{A-A}}$
	B_2	1 mol B–B bonds	1 mol	432 kJ/mol
products	AB	1 mol A–B bonds	2 mol	$D_{\text{A-B}}$

Step 3: Substitute values in the formula for calculating the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$-549 \text{ kJ} = 1 \text{ mol} \times D_{\text{A–A}} + 1 \text{ mol} \times D_{\text{B–B}} - 2 \text{ mol} \times D_{\text{A–B}}$$

$$-549 \text{ kJ} = (1 \text{ mol} \times D_{\text{A–A}}) + \left(1 \cancel{\text{mol}} \times \frac{432 \text{ kJ}}{\cancel{\text{mol}}}\right) - (4 \text{ mol} \times D_{\text{A–A}})$$

$$-549 \text{ kJ} = 432 \text{ kJ} - 3 \text{ mol} \times D_{\text{A–A}}$$

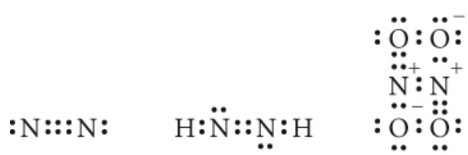
$$3 \text{ mol} \times D_{\text{A–A}} = 432 \text{ kJ} + 549 \text{ kJ}$$

$$D_{\text{A–A}} = \frac{981 \text{ kJ}}{3 \text{ mol}}$$

$$D_{\text{A–A}} = 327 \text{ kJ/mol}$$

Statement: The bond energy of A_2 is 327 kJ/mol.

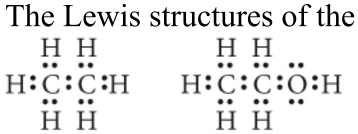
5. The Lewis structures of the molecules are:



N_2 N_2H_2 N_2O_4

Nitrogen gas has the shortest nitrogen–nitrogen bond distance because it contains a nitrogen–nitrogen triple bond. As the number of bonds increases, the bond length shortens.

6. The Lewis structures of the molecules are:



ethane ethanol

The bond energy of a bond is affected by the number and types of atoms and bonds around it. The ethane molecule contains two carbons forming only C–H bonds, while the ethanol molecule containing two carbons has one of the carbons forming a C–O bond. Therefore, the numerical values of a carbon–carbon bond in the two molecules are slightly different.

7. The bond energy of a C–O bond is less than that of a C=O bond because in a C=O bond, more electrons are shared between the two atoms, making the energy required to break the bond greater.

8. Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$CH_3OH(g) + CO(g) \rightarrow CH_3COOH(l)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	CH_3OH	3 mol C–H bonds 1 mol C–O bonds 1 mol H–O bonds	3 mol 1 mol 1 mol	413 kJ/mol 358 kJ/mol 467 kJ/mol
	CO	1 mol C≡O bonds	1 mol	1072 kJ/mol
products	CH_3COOH	3 mol C–H bonds 1 mol C–C bonds 1 mol C=O bonds 1 mol C–O bonds 1 mol H–O bonds	3 mol 1 mol 1 mol 1 mol 1 mol	413 kJ/mol 347 kJ/mol 745 kJ/mol 358 kJ/mol 467 kJ/mol

Step 3: Calculate the enthalpy change, Δ*H*, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= \cancel{3\text{ mol} \times D_{\text{C-H}}} + \cancel{1\text{ mol} \times D_{\text{C-O}}} + \cancel{1\text{ mol} \times D_{\text{H-O}}} + 1\text{ mol} \times D_{\text{C}\equiv\text{C}} -$$

$$\left(\cancel{3\text{ mol} \times D_{\text{C-H}}} + 1\text{ mol} \times D_{\text{C-C}} + 1\text{ mol} \times D_{\text{C=O}} + \cancel{1\text{ mol} \times D_{\text{C-O}}} + \cancel{1\text{ mol} \times D_{\text{H-O}}}\right)$$

$$= \left(1\ \cancel{\text{mol}} \times \frac{1072\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(1\ \cancel{\text{mol}} \times \frac{347\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{745\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 1072\text{ kJ} - (347\text{ kJ} + 745\text{ kJ})$$

$$\Delta H = -20\text{ kJ}$$

Statement: The Δ*H* for the reaction is –20 kJ.

9. Given: $\Delta H = -549$ kJ ; $D_{\text{C-C}} = 347$ kJ/mol ; $D_{\text{C=C}} = 614$ kJ/mol ; $D_{\text{F-F}} = 154$ kJ/mol

Required: bond energy of carbon–fluorine, $D_{\text{C-F}}$

Analysis: $\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$

Solution:

Step 1: Use the balanced chemical equation to determine the bonding of each substance.

$C_2H_4(g) + F_2(g) \rightarrow C_2H_4F_2(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed. Record this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	C_2H_4	4 mol C–H bonds 1 mol C=C bonds	4 mol 1 mol	413 kJ/mol 614 kJ/mol
	F_2	1 mol F–F bonds	1 mol	154 kJ/mol
products	$C_2H_4F_2$	4 mol C–H bonds 1 mol C–C bonds 2 mol C–F bonds	4 mol 1 mol 2 mol	413 kJ/mol 347 kJ/mol $D_{\text{C-F}}$

Step 3: Substitute values in the formula for calculating the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$-549\text{ kJ} = \cancel{4\text{ mol} \times D_{\text{C-H}}} + 1\text{ mol} \times D_{\text{C=C}} + 1\text{ mol} \times D_{\text{F-F}} - \left(\cancel{4\text{ mol} \times D_{\text{C-H}}} + 1\text{ mol} \times D_{\text{C-C}} + 2\text{ mol} \times D_{\text{C-F}}\right)$$

$$-549\text{ kJ} = \left(1\ \cancel{\text{mol}} \times \frac{614\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{154\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(1\ \cancel{\text{mol}} \times \frac{347\text{ kJ}}{\cancel{\text{mol}}}\right) + 2\text{ mol } D_{\text{C-F}}\right]$$

$$-549\text{ kJ} = 614\text{ kJ} + 154\text{ kJ} - 347\text{ kJ} - 2\text{ mol } D_{\text{C-F}}$$

$$2\text{ mol } D_{\text{C-F}} = 614\text{ kJ} + 154\text{ kJ} - 347\text{ kJ} + 549\text{ kJ}$$

$$D_{\text{C-F}} = \frac{970\text{ kJ}}{2\text{ mol}}$$

$$D_{\text{C-F}} = 485\text{ kJ/mol}$$

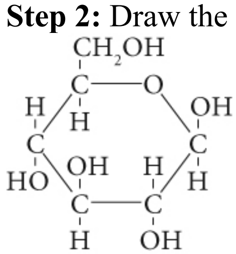
Statement: The carbon–fluorine bond energy is 485 kJ/mol.

10. Solution:

Step 1: Write a balanced chemical equation for the production of glucose.

$6\ CO_2(g) + 6\ H_2O(l) + \text{energy} \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$

Step 2: Draw the structure of glucose to determine the bonding in the substance.



Step 3: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	**Number of bonds per mole ($n_{\text{substance}}$)**	**Amount of bonds in reaction**	**Bond energy per mole**
reactants	CO_2	2 mol C=O bonds	12 mol	799 kJ/mol
	H_2O	2 mol O–H bonds	12 mol	467 kJ/mol
products	$C_6H_{12}O_6$	7 mol C–H bonds 5 mol C–C bonds 5 mol O–H bonds 7 mol C–O bonds	7 mol 5 mol 5 mol 7 mol	413 kJ/mol 347 kJ/mol 467 kJ/mol 358 kJ/mol
	O_2	1 mol O=O bonds	6 mol	495 kJ/mol

Step 4: Calculate the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 12\text{ mol } \times D_{\text{C=O}} + 12\text{ mol} \times\ D_{\text{H–O}} -$$

$$\left(7\text{ mol} \times D_{\text{C–H}} + 5\text{ mol} \times D_{\text{C–C}} + 5\text{ mol} \times D_{\text{H–O}} + 7\text{ mol} \times D_{\text{C–O}} + 6\text{ mol} \times D_{\text{O=O}}\right)$$

$$= \left(12\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(12\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(7\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(5\ \cancel{\text{mol}} \times \frac{347\text{ kJ}}{\cancel{\text{mol}}}\right) +\right.$$

$$\left.\left(5\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(7\ \cancel{\text{mol}} \times \frac{358\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(6\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 9588\text{ kJ} + 5604\text{ kJ} - (2891\text{ kJ} + 1735\text{ kJ} + 2335\text{ kJ} + 2506\text{ kJ} + 2970\text{ kJ})$$

$$\Delta H = 2755\text{ kJ}$$

Statement: The amount of energy the Sun provides to produce 1 mol of glucose is estimated to be 2755 kJ.

11. Solution:

Step 1: Write a balanced chemical equation for the reaction.

$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$

Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	CO	1 mol C≡O bonds	2 mol	1072 kJ/mol
	O_2	1 mol O=O bonds	1 mol	495 kJ/mol
products	CO_2	2 mol C=O bonds	4 mol	799 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 2\text{ mol} \times D_{\text{C≡O}} + 1\text{ mol} \times D_{\text{O=O}} - 4\text{ mol} \times D_{\text{C=O}}$$

$$= \left(2\ \cancel{\text{mol}} \times \frac{1072\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right) - \left(4\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right)$$

$$= 2144\text{ kJ} + 495\text{ kJ} - 3196\text{ kJ}$$

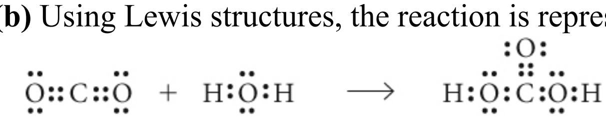
$$\Delta H = -557\text{ kJ}$$

Statement: Since ΔH is negative, the reaction of carbon monoxide with oxygen to produce carbon dioxide is an exothermic reaction.

12. (a) A balanced chemical equation for the reaction is:

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

(b) Using Lewis structures, the reaction is represented as:



CO_2 H_2O H_2CO_3

(c) Solution:
Step 1: Use the Lewis structures to determine the bonding of each substance.
Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	CO_2	2 mol C=O bonds	2 mol	799 kJ/mol
	H_2O	2 mol H–O bonds	2 mol	467 kJ/mol
products	H_2CO_3	1 mol C=O bonds 2 mol C–O bonds 2 mol O–H bonds	1 mol 2 mol 2 mol	745 kJ/mol 358 kJ/mol 467 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 2\text{ mol} \times D_{\text{C=O}} + \cancel{2\text{ mol} \times D_{\text{H-O}}} - (1\text{ mol} \times D_{\text{C=O}} + 2\text{ mol} \times D_{\text{C-O}} + \cancel{2\text{ mol} \times D_{\text{H-O}}})$$

$$= \left(2\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(1\ \cancel{\text{mol}} \times \frac{745\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{358\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 1598\text{ kJ} - (745\text{ kJ} + 716\text{ kJ})$$

$$\Delta H = 137\text{ kJ}$$

Statement: The energy required to convert 1 mol of carbon dioxide to 1 mol of carbonic acid is 137 kJ.

13. (a) Solution:
Step 1: Write the equation that represents the complete combustion of propane.
$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$
Write the equation that represents the complete combustion of methane.
$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$
Step 2: Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.
Combustion of propane:

	Substance	Number of bonds per mole ($n_{\text{substance}}$)	Amount of bonds in reaction	Bond energy per mole
reactants	C_3H_8	8 mol C–H bonds 2 mol C–C bonds	8 mol 2 mol	413 kJ/mol 347 kJ/mol
	O_2	1 mol O=O bonds	5 mol	495 kJ/mol
products	CO_2	2 mol C=O bonds	6 mol	799 kJ/mol
	H_2O	2 mol O–H bonds	8 mol	467 kJ/mol

Combustion of methane:

	Substance	Number of bonds per mole ($n_{substance}$)	Amount of bonds in reaction	Bond energy per mole
reactants	CH_4	4 mol C–H bonds	4 mol	413 kJ/mol
	O_2	1 mol O=O bonds	2 mol	495 kJ/mol
products	CO_2	2 mol C=O bonds	2 mol	799 kJ/mol
	H_2O	2 mol H–O bonds	4 mol	467 kJ/mol

Step 3: Calculate the enthalpy change, ΔH, of each combustion reaction.

Combustion of propane:

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 8\text{ mol} \times D_{\text{C-H}} + 2\text{ mol} \times D_{\text{C-C}} + 5\text{ mol} \times D_{\text{O=O}} - \left(6\text{ mol} \times D_{\text{C=O}} + 8\text{ mol} \times D_{\text{H-O}}\right)$$

$$= \left(8\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{347\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(5\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(6\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(8\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 3304\text{ kJ} + 694\text{ kJ} + 2475\text{ kJ} - (4794\text{ kJ} + 3736\text{ kJ})$$

$$\Delta H = -2057\text{ kJ}$$

Combustion of methane:

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 4\text{ mol} \times D_{\text{C-H}} + 2\text{ mol} \times D_{\text{O=O}} - \left(2\text{ mol} \times D_{\text{C=O}} + 4\text{ mol} \times D_{\text{H-O}}\right)$$

$$= \left(4\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{495\text{ kJ}}{\cancel{\text{mol}}}\right) - \left[\left(2\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(4\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 1652\text{ kJ} + 990\text{ kJ} - (1598\text{ kJ} + 1868\text{ kJ})$$

$$\Delta H = -824\text{ kJ}$$

Statement: The energy produced by the complete combustion of 1 mol of propane is 2057 kJ and that produced by 1 mol of methane is 824 kJ.

(b) Divide the amount of energy produced by 1.0 mol propane by the amount of energy produced by 1.0 mol methane (natural gas).

$$\frac{2057\text{ kJ}}{824\text{ kJ}} = 2.5$$

To produce the same amount of energy as 1.0 mol of propane, 2.5 mol of natural gas is needed.