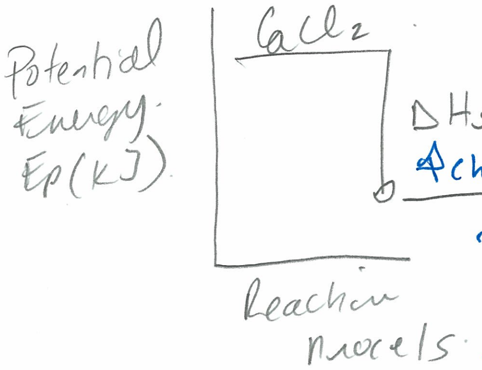
Answers to Review Question

Bond energies can be used to calculate enthalpy changes

Draw and annotate potential energy diagrams for exothermic and endothermic reactions. Identify endothermic bond breaking, exothermic new bond formation, and the enthalpy change. Use the energies of the bonds to write an expression for the enthalpy change.

Cl - Ca - Cl broken energy released.



$\Delta H_{sol} = -82.8$

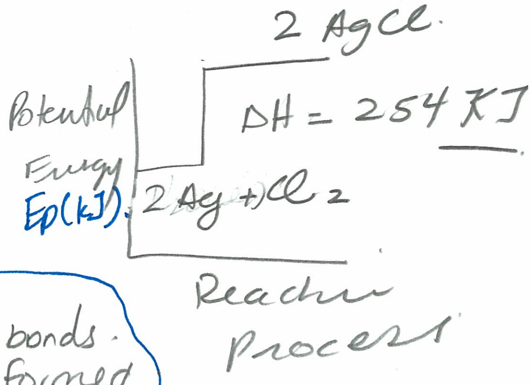
Δ change in Enthalpy.

$$\boxed{\Delta H = \Sigma n D_{\text{bonds broken}} - \Sigma n D_{\text{bonds formed}}}$$

the energy released to or absorbed from the surroundings during a chemical or physical change.

(-) exothermic. ΔH.

(+) endothermic ΔH



Check out the table of bond dissociation energies, p.307 table 1.

C-H	413 .	C-C	347 kJ/mol.
C-O	358	C=C	614
C=O	745 not CO2	C≡C	839
O-H	467	C=O	799 kJ — for carbon dioxide

How much energy is released when 2 moles of hydrogen gas are mixed with 1 mole of fluorine gas in a sealed container and allowed to react?

2 H - H + 1 :F:F:

H - H.

$$2 \text{ mole} \times \frac{432 \text{ kJ}}{\text{mol}} + 1 \text{ mol} \times \frac{154 \text{ kJ}}{\text{mol}}$$

=

$= 864 \text{ kJ} + 154 \text{ kJ}$

$= 1018 \text{ kJ}$.

Methane can be reacted with chlorine and fluorine to form freon-12, CCl_2F_2, hydrogen fluoride, and hydrogen chloride. Calculate the enthalpy change for the reaction.

p.313 #1,2,11,13

H_4C (H-C(H)(H)-H) + Cl_2 + F_2 → Cl-C(F)(Cl)-F + 2 H_2.

Bonds broken - Bonds formed.

$$4 \text{ mol}_{CH} \times \frac{\text{kJ}}{\text{mo}} + 2 \text{ mol}_{Cl-Cl} \times \frac{\text{kJ}}{\text{mol}} + 1 \text{ mol}_{F-F} \times \frac{\text{kJ}}{\text{mole}}.$$

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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, ΔH, of the reaction.

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

$$= 1\ \text{mol} \times D_{\text{H–H}} + 1\ \text{mol} \times D_{\text{Cl–Cl}} - 2\ \text{mol} \times D_{\text{H–Cl}}$$

$$= \left(1\ \cancel{\text{mol}} \times \frac{432\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{239\ \text{kJ}}{\cancel{\text{mol}}}\right) - \left(2\ \cancel{\text{mol}} \times \frac{427\ \text{kJ}}{\cancel{\text{mol}}}\right)$$

$$= 432\ \text{kJ} + 239\ \text{kJ} - 854\ \text{kJ}$$

$$\Delta H = -183\ \text{kJ}$$

Statement: The ΔH 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, ΔH, of the reaction.

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

$$= 1\ \text{mol} \times D_{\text{N≡N}} + 1\ \text{mol} \times D_{\text{H–H}} - 2\ \text{mol} \times D_{\text{N–H}}$$

$$= \left(1\ \cancel{\text{mol}} \times \frac{941\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(3\ \cancel{\text{mol}} \times \frac{432\ \text{kJ}}{\cancel{\text{mol}}}\right) - \left(6\ \cancel{\text{mol}} \times \frac{391\ \text{kJ}}{\cancel{\text{mol}}}\right)$$

$$= 941\ \text{kJ} + 1296\ \text{kJ} - 2346\ \text{kJ}$$

$$\Delta H = -109\ \text{kJ}$$

Statement: The ΔH 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, ΔH, of the reaction.

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

$$= 1\ \text{mol} \times D_{\text{C–H}} + 1\ \text{mol} \times D_{\text{C≡N}} + 2\ \text{mol} \times D_{\text{H–H}} -$$

$$(3\ \text{mol} \times D_{\text{C–H}} + 1\ \text{mol} \times D_{\text{C–N}} + 2\ \text{mol} \times D_{\text{N–H}})$$

$$= \left(1\ \cancel{\text{mol}} \times \frac{413\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{891\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{432\ \text{kJ}}{\cancel{\text{mol}}}\right) -$$

$$\left[\left(3\ \cancel{\text{mol}} \times \frac{413\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(1\ \cancel{\text{mol}} \times \frac{305\ \text{kJ}}{\cancel{\text{mol}}}\right) + \left(2\ \cancel{\text{mol}} \times \frac{391\ \text{kJ}}{\cancel{\text{mol}}}\right)\right]$$

$$= 413\ \text{kJ} + 891\ \text{kJ} + 864\ \text{kJ} - (1239\ \text{kJ} + 305\ \text{kJ} + 782\ \text{kJ})$$

$\Delta H = -158$ kJ

Statement: The ΔH 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 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.

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

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

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

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

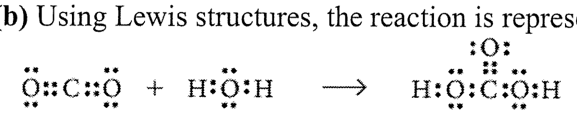
$$\Delta H = -557 \text{ kJ}$$

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 ($n_{substance}$)	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	1 mol	745 kJ/mol
		2 mol C–O bonds	2 mol	358 kJ/mol
		2 mol O–H bonds	2 mol	467 kJ/mol

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

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

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

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

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

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

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 ($n_{substance}$)	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

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:

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

$$= 8\text{ mol} \times D_{\text{C-H}} + 2\text{ mol} \times D_{\text{C-C}} + 5\text{ mol} \times D_{\text{O=O}} -$$

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

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

$$\left[\left(6\ \cancel{\text{mol}} \times \frac{799\text{ kJ}}{\cancel{\text{mol}}}\right) + \left(8\ \cancel{\text{mol}} \times \frac{467\text{ kJ}}{\cancel{\text{mol}}}\right)\right]$$

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

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

Combustion of methane:

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

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

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

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

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

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

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.