

Announcements for Monday, 18NOV2024

- **RE-OPENED:** Week 10 Homework Assignments available on eLearning
 - Graded and Timed Quiz 10 – “Reactions in aqueous solution” due **Wednesday, 20NOV2024, 6:00 PM (EST)**
- Replacement Grade Policy
 - See Canvas announcement from Nov 14 for details

ANY GENERAL QUESTIONS? Feel free to see me after class!

Try This On Your Own

- The enthalpy of combustion for 1 mole of propane (C_3H_8) is -2044 kJ . What is the minimum mass of propane needed to react in excess oxygen to generate enough heat to raise the temperature of 10.0 moles of silver by 25.0°C ? The specific heat of silver is $0.235 \text{ J/g}\cdot^\circ\text{C}$. **$0.137 \text{ g C}_3\text{H}_8$**

$$10.0 \text{ mol Ag} \times \frac{107.87 \text{ g}}{1 \text{ mol Ag}} = 1079 \text{ g Ag}$$

heated needed to raise the temperature of 10.0 mol silver by 25.0°C

$$q_{\text{Ag}} = m_{\text{Ag}} \times C_{s\text{Ag}} \times \Delta T_{\text{Ag}} = (1079 \text{ g})(0.235 \text{ J/g}\cdot^\circ\text{C})(25.0^\circ\text{C}) = 6339 \text{ J or } 6.339 \text{ kJ absorbed}$$

$$q_{\text{Ag}} = -q_{\text{rxn}} \rightarrow q_{\text{rxn}} = -6.339 \text{ kJ}$$

mass of C_3H_8 needed to release 6.339 kJ

$$-6.339 \text{ kJ} \times \frac{1 \text{ mol C}_3\text{H}_8}{-2044 \text{ kJ}} \times \frac{44.09 \text{ g}}{1 \text{ mol C}_3\text{H}_8} = \mathbf{0.137 \text{ g C}_3\text{H}_8}$$

Try This On Your Own

250.0 mL of 1.00 M HCl(aq) was added to 250.0 mL of 1.00 M NaOH(aq) in a coffee-cup calorimeter. Both solutions initially are at 20.00 °C. Upon mixing, the temperature of the mixture increases to 26.84 °C. Using this data, calculate the enthalpy change for the following reaction:



Assume that the density and heat capacity of the mixture are 1.00 g/mL and 4.18 J/g·°C, respectively.

Some questions to consider:

Why is the temperature of the mixture increasing? **It is absorbing heat**

Where is the heat that warms the mixture coming from? **From the reaction**

What can we deduce about the sign of ΔH_{rxn} ? **ΔH_{rxn} is (-) , exothermic**

How does ΔH_{rxn} relate to q_{solution} ? **$\Delta H_{\text{rxn}} = -q_{\text{solution}}$**

What amounts of HCl and NaOH are actually reacting? **0.250 moles of HCl and NaOH**

What amounts of HCl and NaOH are being depicted in the reaction? **1 mole of each**

Try This On Your Own

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Assume that the density and heat capacity of the mixture are 1.00 g/mL and 4.18 J/g·°C, respectively.

250.0 mL of 1.00 M HCl = 0.250 mol HCl

250.0 mL of 1.00 M NaOH = 0.250 mol NaOH

Total volume of solution = 500.0 mL = 500. g

$\Delta H_{\text{rxn}} = -q_{\text{solution}}$ where $q_{\text{solution}} = mC_s\Delta T = (500. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(26.84^\circ\text{C} - 20.00^\circ\text{C})$

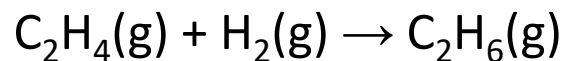
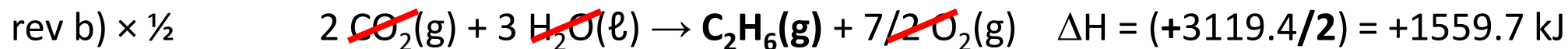
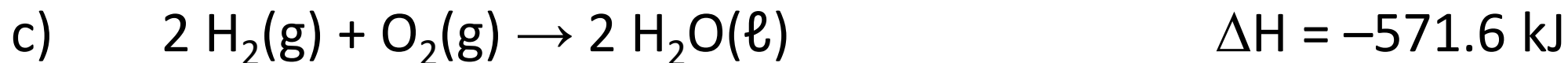
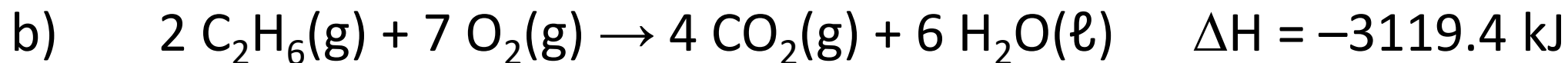
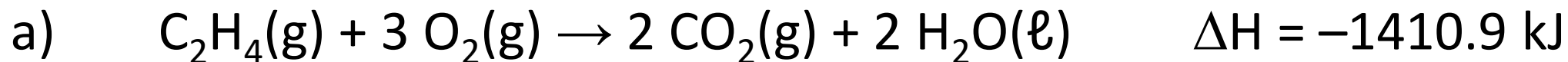
$q_{\text{solution}} = 14295.6 \text{ J} = 14.2956 \text{ kJ}$, $\Delta H_{\text{rxn}} = -14.2956 \text{ kJ}/0.250 \text{ mol} = \mathbf{-57.2 \text{ kJ}}$

Try This On Your Own

Calculate the enthalpy change for the following reaction



given the following data



$$\Delta H = (-1410.9) + (1559.7) + (-285.8) = -137.0 \text{ kJ}$$

Estimating ΔH_{rxn} from Bond Energies

- in chemical reactions some/all existing bonds are broken and new bonds are made
 - breaking bonds **requires** energy (i.e., is endothermic; $\Delta H (+)$)
 - making bonds **releases** energy (i.e., is exothermic; $\Delta H (-)$)
- Recall from Chapter 5: **average bond energy** = amount of energy needed to break 1 mole of bonds in the gas phase

TABLE 9.3 Average Bond Energies

Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)
H—H	436	N—N	163	Br—F	237
H—C	414	N=N	418	Br—Cl	218
H—N	389	N≡N	946	Br—Br	193
H—O	464	N—O	222	I—Cl	208

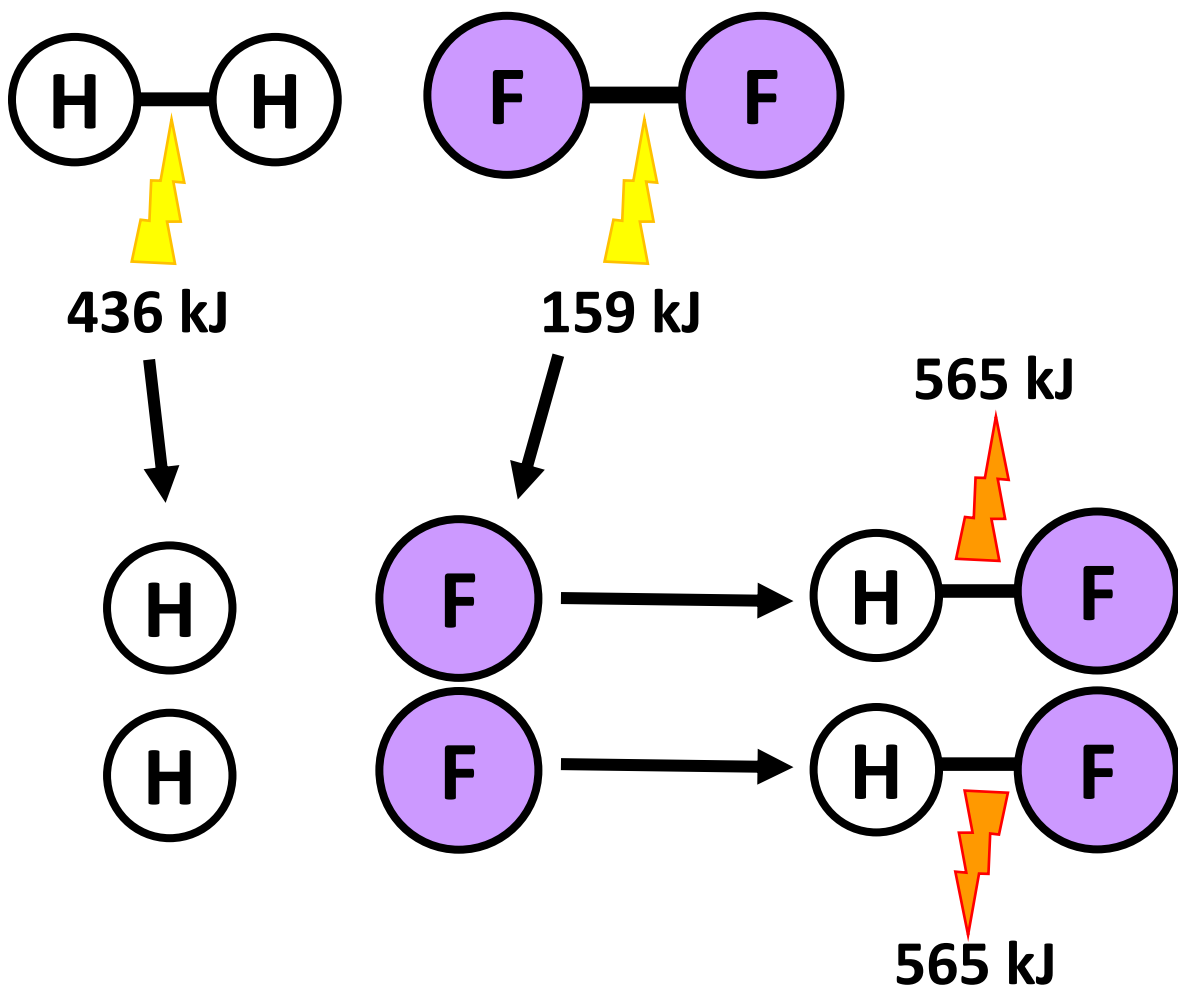
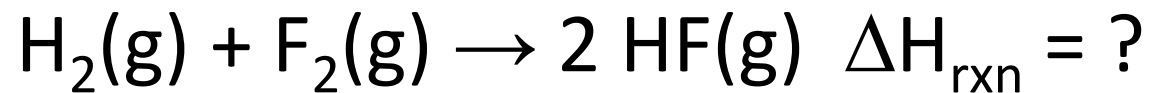
Estimating ΔH_{rxn} from Bond Energies (continued)

- enthalpy changes for chemical reactions *in the gas phase* can be estimated by summing the amounts of energy **absorbed** to break bonds and the amounts of energy **released** when new bonds are made during the course of a reaction.

$$\Delta H_{\text{rxn}} = \underbrace{\sum(\Delta H\text{'s bonds broken})}_{\text{Positive}} + \underbrace{\sum(\Delta H\text{'s bonds formed})}_{\text{Negative}}$$

Estimating ΔH_{rxn} from Bond Energies (continued)

Estimate the enthalpy change for



Bond	Bond Energy (kJ/mol)
H—H	436
F—F	159
H—F	565

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Bonds Broken

$$1 \text{ mol H—H} = +436 \text{ kJ}$$

$$1 \text{ mol F—F} = +159 \text{ kJ}$$

$$+595 \text{ kJ}$$

Bonds Made

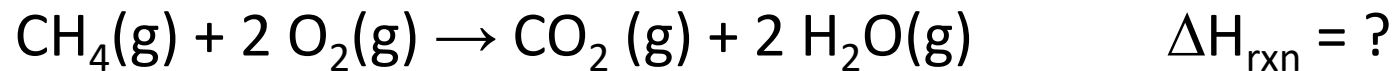
$$2 \text{ mol H—F} = (2)(-565 \text{ kJ})$$

$$-1130 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = 595 \text{ kJ} + (-1130 \text{ kJ}) = \mathbf{-535 \text{ kJ}}$$

Try This

Estimate the enthalpy change for the complete combustion of 1 mole of methane



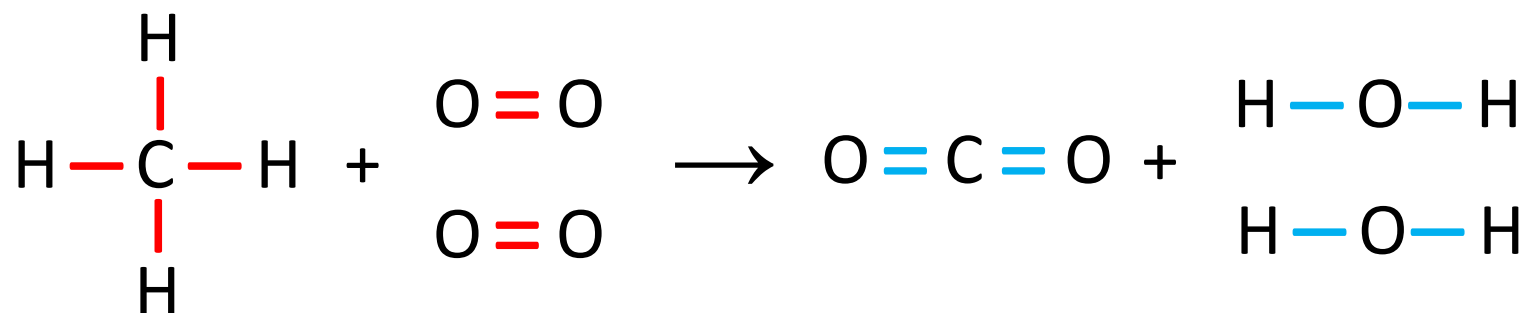
given the following average bond energies:

$$\text{H}-\text{C} = 414 \text{ kJ/mol}$$

$$\text{H}-\text{O} = 464 \text{ kJ/mol}$$

$$\text{C}=\text{O} \text{ in } \text{CO}_2 = 799 \text{ kJ/mol}$$

$$\text{O}=\text{O} = 498 \text{ kJ/mol}$$



Bonds Broken

$$4 \text{ moles H}-\text{C} = (4)(414) = \mathbf{1656 \text{ kJ}}$$

$$2 \text{ moles O}=\text{O} = (2)(498) = \mathbf{996 \text{ kJ}}$$
$$\mathbf{+2652 \text{ kJ}}$$

Bonds Made

$$4 \text{ moles H}-\text{O} = (4)(-464) = \mathbf{-1856 \text{ kJ}}$$

$$2 \text{ moles C}=\text{O} = (2)(-799) = \mathbf{-1598 \text{ kJ}}$$
$$\mathbf{-3454 \text{ kJ}}$$

$$\Delta H_{\text{rxn}} = \Sigma(\Delta H \text{ bonds broken}) + \Sigma(\Delta H \text{ bonds made}) = \mathbf{2652 \text{ kJ}} + (\mathbf{-3454 \text{ kJ}}) = \mathbf{-802 \text{ kJ}}$$

Determining ΔH_{rxn} from Standard Enthalpies of Formation

First: What exactly is a Standard Enthalpy of Formation (ΔH_f°) for a Compound?

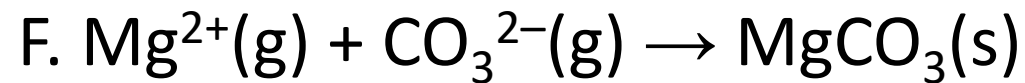
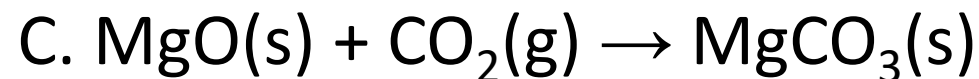
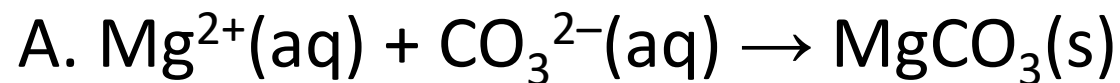
- it's an enthalpy change for a **VERY SPECIFIC reaction**
 - **ONLY 1 mole** of compound is formed
 - the compound must be formed **from its constituent elements**
 - the elements must be **in their standard states (!?!)**

standard state of a substance:

- **for a Gas:** the standard state is the pure gas at a pressure of exactly 1 atmosphere (atm)
- **for a Liquid or Solid:** the standard state is the pure substance *in its most stable form* at a pressure of 1 atm and at the temperature of interest (usually 25 °C)
- **for a Substance in Solution:** the standard state is at a concentration of exactly 1 M
- When a reaction takes place between reactants **in their standard states** to form products **in their standard states**, the resulting enthalpy change is a **STANDARD ENTHALPY CHANGE (ΔH°)**

Try This

ΔH_f° for magnesium carbonate is -1095.8 kJ/mol. To which of the following chemical equations does this enthalpy change apply?



Beware! You MUST be able to write the proper chemical equation when given a value for ΔH_f°

Determining ΔH_{rxn} from Standard Enthalpies of Formation (continued)

- Second: How can the enthalpy change of a reaction can be calculated from tabulated standard enthalpies of formation?

- note that the enthalpies of formation for pure **elements in their standard state** are assigned values of zero!

Standard Enthalpies (or Heats) of Formation, ΔH_f° , at 298 K

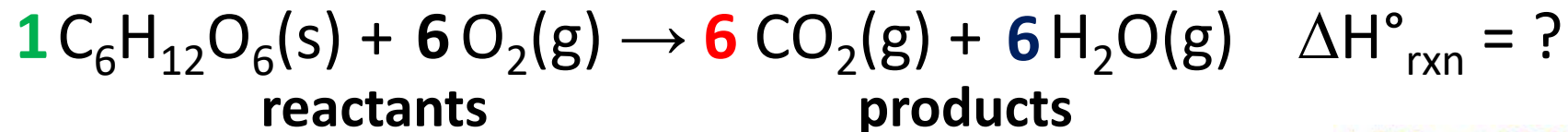
Formula	ΔH_f° (kJ/mol)	Formula	ΔH_f° (kJ/mol)
Bromine		$\text{C}_3\text{H}_8\text{O}(l, \text{isopropanol})$	-318.1
$\text{Br}(g)$	111.9	$\text{C}_6\text{H}_6(l)$	49.1
$\text{Br}_2(l)$	0	$\text{C}_6\text{H}_{12}\text{O}_6(s, \text{glucose})$	-1273.3
$\text{HBr}(g)$	-36.3	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(s, \text{sucrose})$	-2226.1
Calcium		Chlorine	
$\text{Ca}(s)$	0	$\text{Cl}(g)$	121.3
$\text{CaO}(s)$	-634.9	$\text{Cl}_2(g)$	0
$\text{CaCO}_3(s)$	-1207.6	$\text{HCl}(g)$	-92.3
Carbon		Fluorine	
$\text{C}(s, \text{graphite})$	0	$\text{F}(g)$	79.38
$\text{C}(s, \text{diamond})$	1.88	$\text{F}_2(g)$	0

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$$\Delta H_{\text{rxn}}^\circ = \sum n_p \Delta H_f^\circ (\text{products}) - \sum n_r \Delta H_f^\circ (\text{reactants})$$

- where n_p and n_r = stoichiometric coefficients of the products and reactants, respectively
- use parentheses and brackets so that you don't lose track of the minus signs!

Use standard enthalpies of formation to determine the standard enthalpy change for the complete combustion of 1 mole of glucose:



ΔH°_f , at 298 K

$$\Delta H^\circ_{\text{rxn}} = \sum n_p \Delta H^\circ_f (\text{products}) - \sum n_r \Delta H^\circ_f (\text{reactants})$$

$$\Delta H^\circ_{\text{rxn}} = \left[\textcolor{red}{(6)(-393.5)} + \textcolor{blue}{(6)(-241.8)} \right] - \left[\textcolor{teal}{(1)(-1273.3)} + (6)(0) \right]$$

$$\left[-3811.8 \right] - \left[-1273.3 \right]$$

$$-3811.8 + 1273.3$$

$$\Delta H^\circ_{\text{rxn}} = \textcolor{violet}{-2538.5 \text{ kJ/mol}}$$

Formula	ΔH°_f (kJ/mol)
CO ₂ (g)	-393.5
C ₆ H ₁₂ O ₆ (s)	-1273.3
H ₂ O(g)	-241.8