Unit 4: Chemical Quantities and Thermochemistry

Chapter 7: Chemical Quantities

7.1: The Mole: A Measurement of Matter

Mole (mol): - a group of atoms or molecules numbered **6.02×10^{23} (Avogadro's Number)**

Example: 1 mol of carbon (C) = 6.02×10^{23} carbon atoms = 12.01 g (same as the amu)
        1 mol of oxygen (O_2) = 6.02×10^{23} oxygen molecules = 32.00 g (include subscripts with amu)

Example 1: Calculate the mass of 250 atoms of gold.

1 mol Au = 196.97 g Au = 6.02×10^{23} Au atoms

$$250 \text{ atoms} \times \frac{196.97 \text{ g}}{6.02 \times 10^{23} \text{ atoms}}$$

8.18×10^{-20} g of Au

Example 2: Determine the number of molecules for 50.0 mg of oxygen.

1 mol O_2 = 32.00 g O_2 = 6.02×10^{23} O_2 molecules
(16.00 × 2)

50.0 mg = 0.0500 g

$$0.0500 \text{ g} \times \frac{6.02 \times 10^{23} \text{ molecules}}{32.00 \text{ g}}$$

9.41×10^{20} molecules of O_2

Molar Mass (g/mol): - the mass per one mole of atoms, molecules or formula unit.
- also called **gram atomic mass (gam)**, **gram molecular mass (gmm) for molecular compound**, or **gram formula mass (gfm) for ionic compound**.
- molar mass of a mono-atomic element is the same as the atomic mass (gam).
- molar mass of a compound, binary element, or polyatomic element is the same as the combine atomic masses of all atoms in the molecule (gmm).
- molar mass of an ionic compound is the same as the combine atomic masses of all atoms in the ionic formula unit (gfm).

Example 3: Find the molar mass of the following.

a. aluminium

Al = 26.98 g/mol

b. nitrogen

N_2 = 2 × 14.01 g/mol

N_2 = 28.02 g/mol

c. water

H_2O = (2 × 1.01) + 16.00

H_2O = 18.02 g/mol

d. lead (IV) nitrate

$Pb(NO_3)_4$ = 207.21 + (4 × 14.01) + (12 × 16.00)

$Pb(NO_3)_4$ = 455.25 g/mol

e. sucrose

$C_{12}H_{22}O_{11}$ = (12 × 12.01) + (22 × 1.01) + (11 × 16.00)

$C_{12}H_{22}O_{11}$ = 342.34 g/mol

7.2: Mole-Mass and Mole-Volume Relationships

Mole-Mass Relationship

Converting between Mass and Moles:

$$\text{Moles (mol)} = \frac{\text{Mass (g)}}{\text{Molar Mass (g/mol)}} \qquad n = \frac{m}{M}$$

n = **moles**   *m* = **mass**   *M* = **Molar mass**

Example 1: Calculate the number of moles for:

a.   20.0 g of magnesium chloride

$MgCl_2 = 24.31 + 2(35.45)$ $M = 95.21$ g/mol
$n = \dfrac{m}{M} = \dfrac{20.0\,g}{95.21\,g/mol}$ $n = 0.210$ mol

b.   4.52 mg of glucose

$C_6H_{12}O_6 = 6(12.01) + 12(1.01) + 6(16.00)$
$M = 180.18$ g/mol
$n = \dfrac{m}{M} = \dfrac{4.52\,mg}{180.18\,g/mol}$ $n = 0.0251$ mmol

Example 2: Determine the mass of the following amount.

a.   8.52 mol of ozone

$O_3 = 3(16.00)$ $M = 48.00$ g/mol
$n = \dfrac{m}{M}$
$m = nM = (8.52\ \text{mol})(48.00\ \text{g/mol})$

$m = 409.\ g$

b.   24.7 mmol of phosphoric acid

$H_3PO_4 = 3(1.01) + 30.97 + 4(16.00)$ $M = 98.00$ g/mol
$n = \dfrac{m}{M}$
$m = nM = (24.7\ \text{mmol})(98.00\ \text{g/mol}) = 2420.6$ mg

$m = 2.42 \times 10^3\ mg = 2.42\ g$

Mole-Volume Relationship

Standard Temperature and Pressure (STP): - the amount of any gas at 0°C and 101.325 kPa (Earth's atmospheric pressure at sea level).

STP = 22.4 L/mol @ 0°C and 101.325 kPa (1 atm)

Note: The amount of gas is determined by temperature, pressure and volume. The type of gas particles has no effect on these variables. (*Avogadro's Hypothesis*)

Example 3: Determine the amount of oxygen gas in a 5.00 L container under STP.

$STP = 22.4$ L/mol
$n = 5.00\ \text{L} \times \dfrac{1\,\text{mol}}{22.4\ \text{L}}$ $n = 0.223$ mol

Example 4: Determine the volume of 3.50 g of nitrogen gas under STP.

STP = 22.4 L/mol
m = 3.50 g N_2
M = 28.02 g/mol
n = ? mol
V = ? L

$$n = \frac{m}{M} = \frac{3.50\text{ g}}{28.02\text{ g/mol}} = 0.124910778\text{ mol}$$

$$V = (0.124910778\cancel{\text{mol}})(22.4\text{ L/}\cancel{\text{mol}}) \qquad \boxed{V = 2.80\text{ L}}$$

Example 5: Calculate the molar mass of a gaseous compound containing carbon and hydrogen if its density is 0.714 g/L at STP.

STP = 22.4 L/mol
D = 0.714 g/L
M = ? (g/mol)

$M = D \times$ (STP)
$M = (0.714$ g/$\cancel{\text{L}}$) $(22.4$ $\cancel{\text{L}}$/mol)
$M = 15.9936$ g/mol
$\boxed{M = 16.0\text{ g/mol}}$

Example 6: Calculate the density in g/L of a sulfur dioxide gas at STP.

STP = 22.4 L/mol
M = 64.06 g/mol (SO_2)
D = ? (g/L)

$M = D \times$ (STP)

$$D = \frac{M}{(\text{STP})} = \frac{(64.06\text{ g/}\cancel{\text{mol}})}{(22.4\text{ L/}\cancel{\text{mol}})} = 2.859821429\text{ g/L} \qquad \boxed{D = 2.86\text{ g/L}}$$

7.3: Percent Composition and Chemical Formulas

Mass Percent: - the mass percentage of each element in a compound.

For Compound $A_xB_yC_z$ with its **Total Mass (m)**, the **Mass Percentages** are:

$$\%A = \frac{m_A}{m} \times 100\% \qquad \%B = \frac{m_B}{m} \times 100\% \qquad \%C = \frac{m_C}{m} \times 100\%$$

For Compound $A_xB_yC_z$ with its **Molar Mass (M)**, the **Mass Percentages** are:

$$\%A = \frac{(x)(M_A)}{M} \times 100\% \qquad \%B = \frac{(y)(M_B)}{M} \times 100\% \qquad \%C = \frac{(z)(M_C)}{M} \times 100\%$$

Example 1: Propane consists of hydrogen and carbon. If 11.36 g of hydrogen is used to form 62.00 g propane, determine the mass percentages of both hydrogen and carbon.

m_{Total} = 62.00 g
m_C = 62.00 g − 11.36 g

m_H = 11.36 g
m_C = 50.64 g

$$\%H = \frac{m_H}{m_{\text{Total}}} \times 100\%$$

$$\%H = \frac{11.36\text{ g}}{62.00\text{ g}} \times 100\%$$

$$\boxed{\%H = 18.32\ \%}$$

$$\%C = \frac{m_C}{m_{\text{Total}}} \times 100\%$$

$$\%C = \frac{50.64\text{ g}}{62.00\text{ g}} \times 100\%$$

$$\boxed{\%C = 81.68\%}$$

Example 2: Chromium (II) nitrate has a mass percentage of 29.5% chromium and 15.9% nitrogen. What are the masses for each element in 28.0 g of chromium (II) nitrate?

Chromium (II) nitrate: $Cr(NO_3)_2$ $m_{Total} = 28.0$ g   **Cr = 29.5%**   **N = 15.9%**

$O = 100\% - 29.5\% - 15.9\%$ **O = 54.6%**

$$\frac{m_{Cr}}{m_{Total}} \times 100\% = \%Cr$$

$$m_{Cr} = \frac{\%Cr}{100\%} \times m_{Total}$$

$$m_{Cr} = 0.295 \times 28.0 \text{ g}$$

$$\boxed{m_{Cr} = 8.26 \text{ g}}$$

$$\frac{m_{N}}{m_{Total}} \times 100\% = \%N$$

$$m_{N} = \frac{\%N}{100\%} \times m_{Total}$$

$$m_{N} = 0.159 \times 28.0 \text{ g}$$

$$\boxed{m_{N} = 4.45 \text{ g}}$$

$$\frac{m_{O}}{m_{Total}} \times 100\% = \%O$$

$$m_{O} = \frac{\%O}{100\%} \times m_{Total}$$

$$m_{O} = 0.546 \times 28.0 \text{ g}$$

$$\boxed{m_{O} = 15.3 \text{ g}}$$

Example 3: Calculate the mass percentage of each element in sodium chromate.

Na_2CrO_4 $M = 161.98$ g/mol

Assume we have 161.98 g (1 mole) of Na_2CrO_4, there are 2 moles of Na, 1 mole of Cr and 4 moles of O:

$$\% \text{ Na} = \frac{(2 \text{ mol})(22.99 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 28.38622052 \%$$

% Na = 28.39 %

$$\% \text{ Cr} = \frac{(1 \text{ mol})(52.00 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 32.10272873 \%$$

% Cr = 32.10 %

$$\% \text{ O} = \frac{(4 \text{ mol})(16.00 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 39.51105075 \%$$

% O = 39.51 %

Empirical Formula: - the simplest ratio between the elements in a chemical formula.

Molecular Formula: - the actual chemical formula of a compound.

> **Molecular Formula = (Empirical Formula)$_n$**
>
> **where *n* = natural number**

Example: $C_6H_{12}O_6$ ⟵ CH_2O
                Molecular Formula for Glucose   Empirical Formula
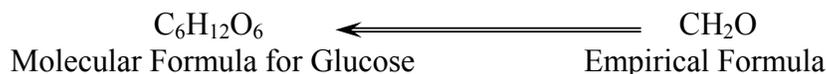
Note: Knowing the mass percentages of a compound allows us to find the empirical formula. To know the molecular formula, we must also know the molar mass.

Example 4: Vitamin C has a molar mass of 176.14 g/mol and contains carbon, hydrogen, and oxygen atoms. If the % mass of carbon and oxygen are 40.91% and 54.50% respectively, determine the empirical and molecular formula of vitamin C.

% C = 40.91%   % O = 54.50%   % H = 100% − 40.91% − 54.50% = 4.59%

Assume 100 g of Vitamin C. Then, there are

$m_C = 100 \text{ g} \times 40.91\% = 40.91 \text{ g}$ $m_O = 100 \text{ g} \times 54.50\% = 54.50 \text{ g}$ $m_H = 100 \text{ g} \times 4.59\% = 4.59 \text{ g}$

$$n_C = \frac{40.91 \text{ g}}{12.011 \text{ g/mol}} = 3.406044459 \text{ mol}_C \qquad n_H = \frac{4.59 \text{ g}}{1.0079 \text{ g/mol}} = 4.554023217 \text{ mol}_H$$

$$n_O = \frac{54.50 \text{ g}}{16.00 \text{ g/mol}} = 3.40625 \text{ mol}_O$$

$$\frac{n_C}{n_O} = \frac{3.406044459 \text{ mol C}}{3.40625 \text{ mol O}} \approx \frac{1 \text{ mol C}}{1 \text{ mol O}} \qquad \frac{n_H}{n_O} = \frac{4.554023217 \text{ mol H}}{3.40625 \text{ mol O}} \approx 1.33 = \frac{4 \text{ mol H}}{3 \text{ mol O}}$$

$n_C : n_O = 1 : 1$ ⟵ **Combine Ratios** ⟶ $n_H : n_O = 4 : 3$

Empirical Formula = $C_3H_4O_3$ **(88.07 g/mol)** $\dfrac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \dfrac{176.14 \text{ g/mol}}{88.07 \text{ g/mol}} = 2$

Molecular Formula = Empirical Formula × 2

Molecular Formula = $C_6H_8O_6$

_____ **OR** _____

Another Method may be used where the Actual Molar Mass becomes the Mass of Vitamin used. Then, the Mole of each Atom is calculated to determine the Molecular Formula first.

$$n_C = \frac{40.91\% \times 176.14 \text{ g}}{12.011 \text{ g/mol}} \approx 6.00 \text{ mol}_C \qquad n_H = \frac{4.59\% \times 176.14 \text{ g}}{1.0079 \text{ g/mol}} \approx 8.00 \text{ mol}_H$$

$$n_O = \frac{54.50\% \times 176.14 \text{ g}}{16.00 \text{ g/mol}} \approx 6.00 \text{ mol}_O \qquad \text{**Molecular Formula ($C_6H_8O_6$) will be found first, then**}$$

the **Empirical Formula ($C_3H_4O_3$) will be stated.**

Assignment

7.1 pg. 174 #3, 4; pg. 175 #5, 6; pg. 179 #7, 8; pg. 181 #9 to 11, 13, 14
7.2 pg. 183 #16 to 19; pg. 184 #20, 21; pg. 185 #22, 23; pg. 186 #24 to 28;
     pg. 348 #31 to 33; pg. 349 #34 to 36; pg. 353 #40 and 41
7.3 pg. 189 #29, 30; pg. 191 #31 to 33; pg. 192 #34; pg. 193 # 35, 36;
     pg. 194 #37, 38; pg. 195 #39 to 43
Ch 7 Review: pg. 198 #45 to 52, 55, 56, 60 to 66

Chapter 9: Stoichiometry

9.1: The Arithmetic of Equations

Stoichiometry: - the calculation of quantities in a chemical reaction.
- the coefficients of various reactants and /or products form mole ratios.
- these mole ratios are for moles, molecules / atoms, and volumes of gases at constant temperatures and pressures (such as STP) of all chemicals involved in the chemical equation.

Example 1: Interpret the chemical equation $4\ NH_{3\ (g)} + 7\ O_{2\ (g)} \rightarrow 4\ NO_{2\ (g)} + 6\ H_2O_{\ (g)}$ in terms of

a. moles.   b. molecules.   c. masses.   d. volumes STP.

	$4\ NH_{3\ (g)}$	$7\ O_{2\ (g)}$	$4\ NO_{2\ (g)}$	$6\ H_2O_{\ (g)}$
a.	**4 moles of NH_3**	**7 moles of O_2**	**4 moles of NO_2**	**6 moles of H_2O**
b.	**4 molecules of NH_3**	**7 molecules of O_2**	**4 molecules of NO_2**	**6 molecules of H_2O**
c.	$m = nM$ $m = (4\ \text{mol})(17.04\ \text{g/mol})$ **$m = 68.16\ g$**	$m = nM$ $m = (7\ \text{mol})(32.00\ \text{g/mol})$ **$m = 224.0\ g$**	$m = nM$ $m = (4\ \text{mol})(46.01\ \text{g/mol})$ **$m = 184.0\ g$**	$m = nM$ $m = (6\ \text{mol})(18.02\ \text{g/mol})$ **$m = 108.1\ g$**
d.	$V = (4\ \text{mol})(22.4\ \text{L/mol})$ **$V = 89.6\ L$**	$V = (7\ \text{mol})(22.4\ \text{L/mol})$ **$V = 157\ L$**	$V = (4\ \text{mol})(22.4\ \text{L/mol})$ **$V = 89.6\ L$**	$V = (6\ \text{mol})(22.4\ \text{L/mol})$ **$V = 134\ L$**

*Note the Law of Conservation of Mass holds after converting mole of each chemical to its mass.

9.2: Chemical Calculations (Gravimetric, STP Stoichiometry)

Mole Ratio: - a ratio between the coefficient of the required chemical amount to the given chemical amount. $\left(\dfrac{\textbf{require coefficient}}{\textbf{given coefficient}} \right)$

Example 1: 4.35 mol of $PCl_{5\ (g)}$ is decomposed into its elements. Write a balance equation and determined the amount of chlorine produced.

$$\underline{4}\ PCl_{5\ (g)} \quad \rightarrow \quad P_{4\ (s)} \quad + \quad \underline{10}\ Cl_{2\ (g)}$$
4.35 mol   ? mol

$$n_{Cl_2} = 4.35\ \text{mol PCl}_5 \times \frac{10\ \text{mol Cl}_2}{4\ \text{mol PCl}_5} = 10.875\ \text{mol Cl}_2 \qquad \boxed{n_{Cl_2} = 10.9\ \text{mol Cl}_2}$$
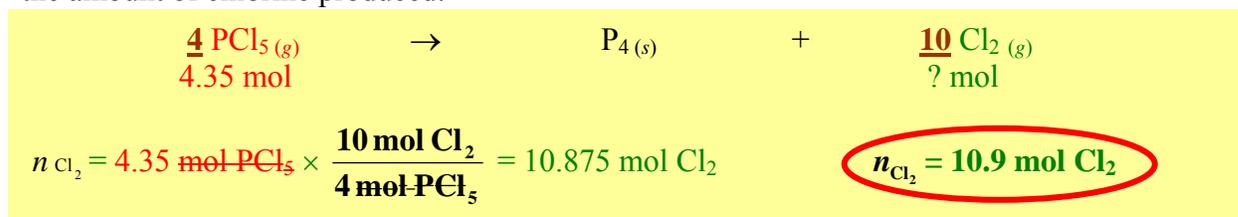
Gravimetric Stoichiometry: - stoichiometry that involves quantities of masses.

Gravimetric Stoichiometry Procedure:

1. Predict the products and balance the chemical equation.

2. Put all the information given under the appropriate chemicals and determine the molar masses of the chemicals involved.

3. Find the moles of the given chemical. $\left(n = \dfrac{m}{M} \right)$

4. Find the mole of the required chemical using mole ratio.

$$\left(\textbf{mol of require} = \textbf{mol of given} \times \dfrac{\textbf{require coefficient}}{\textbf{given coefficient}} \right)$$

5. Convert mole of the required chemical to its mass equivalence. ($m = nM$)

Example 2: Determine the mass of carbon dioxide formed when 50.0 kg of butane ($C_4H_{10\,(l)}$) is burned.

$$2\ C_4H_{10\,(g)} \quad + \quad 13\ O_{2\,(g)} \quad \rightarrow \quad 8\ CO_{2\,(g)} \quad + \quad 10\ H_2O_{\,(g)}$$

50.0 kg   ? g

58.14 g/mol $M = 44.01$ g/mol

$$50.0\ \text{kg } C_4H_{10} \times \frac{1\ \text{mol } C_4H_{10}}{58.14\ \text{g } C_4H_{10}} \times \frac{8\ \text{mol } CO_2}{2\ \text{mol } C_4H_{10}} \times \frac{44.01\ \text{g } CO_2}{1\ \text{mol } CO_2} = 151.3931189\ \text{kg } CO_2$$

$$m_{CO_2} = \textbf{151 kg } CO_2$$
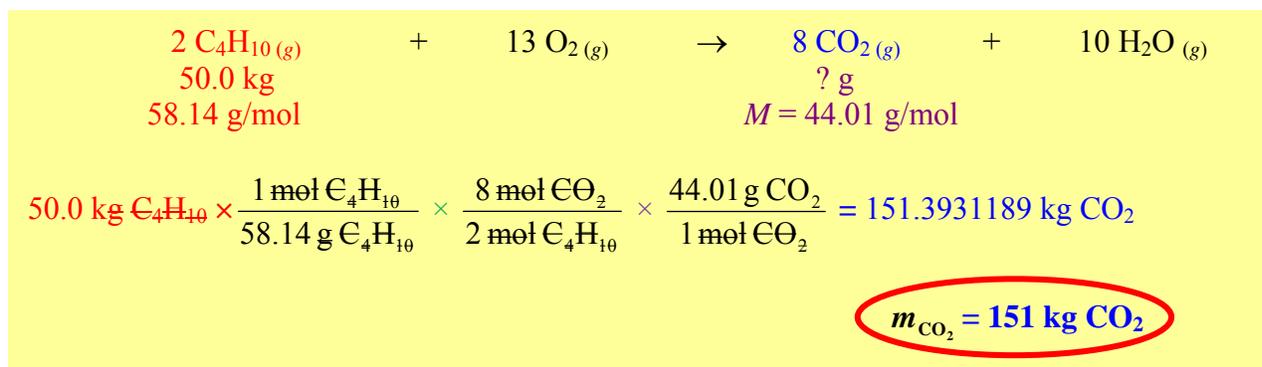
Example 3: Barium bromide solution was mixed with an excess sodium phosphate solution. What was the mass of barium chloride solid needed in the original solution to form 3.21 g of precipitate?
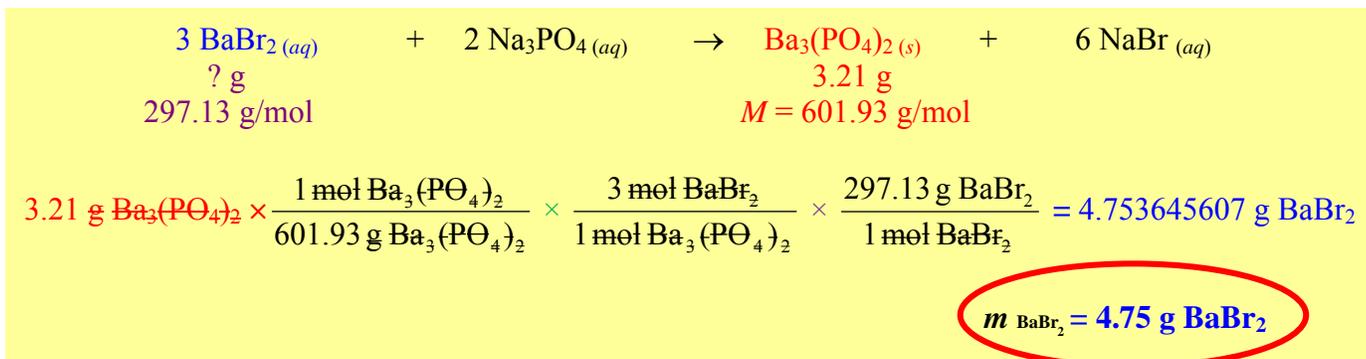
$$3\ BaBr_{2\,(aq)} \quad + \quad 2\ Na_3PO_{4\,(aq)} \quad \rightarrow \quad Ba_3(PO_4)_{2\,(s)} \quad + \quad 6\ NaBr_{\,(aq)}$$

? g   3.21 g

297.13 g/mol $M = 601.93$ g/mol

$$3.21\ \text{g } Ba_3(PO_4)_2 \times \frac{1\ \text{mol } Ba_3(PO_4)_2}{601.93\ \text{g } Ba_3(PO_4)_2} \times \frac{3\ \text{mol } BaBr_2}{1\ \text{mol } Ba_3(PO_4)_2} \times \frac{297.13\ \text{g } BaBr_2}{1\ \text{mol } BaBr_2} = 4.753645607\ \text{g } BaBr_2$$

$$m_{BaBr_2} = \textbf{4.75 g } BaBr_2$$

Gaseous Stoichiometry (STP): - stoichiometry that involves quantities of volumes of gaseous chemicals.

Gaseous Stoichiometry Procedure (STP)

1. Predict the products and balance the chemical equation.

2. Put all the information given under the appropriate chemicals.

3. Find the moles of the given chemical. **(At STP: n = Volume $\times \dfrac{1\,\text{mol}}{22.4\,\text{L}}$)**

4. Find the mole of the required chemical using mole ratio $\left(\textbf{mol of require} = \textbf{mol of given} \times \dfrac{\textbf{require coefficient}}{\textbf{given coefficient}} \right)$

5. Convert mole of the required chemical to its volume equivalence. **(At STP: $V = n \times 22.4$ L/mol)**

Example 4: If 15.25 L of hydrogen at STP is reacted with excess amount of nitrogen, determine the volume of ammonia formed at the same conditions.

$$3\ H_{2\,(g)} \qquad + \qquad N_{2\,(g)} \qquad \rightarrow \qquad 2\ NH_{3\,(g)}$$
$$V = 15.25\ \text{L} \qquad\qquad\qquad\qquad\qquad\quad ?\ \text{L}$$
$$STP = 22.4\ \text{L/mol} \qquad\qquad\qquad\qquad STP = 22.4\ \text{L/mol}$$

$$V_{NH_3} = 15.25\ \cancel{\text{L }H_2} \times \frac{1\ \cancel{\text{mol }H_2}}{22.4\ \cancel{\text{L }H_2}} \times \frac{2\ \text{mol }NH_3}{3\ \cancel{\text{mol }H_2}} \times \frac{22.4\ \text{L }NH_3}{1\ \cancel{\text{mol }NH_3}} = 10.1666667\ \text{L }NH_3$$

$$\boxed{V_{NH_3} = 10.2\ \textbf{L } NH_3}$$

Note: Because the gases are under the same conditions, we could simply do mole ratio only.
(Avogadro's Law)

$$V_{NH_3} = 15.25\ \text{L }\cancel{H_2} \times \frac{2\ \cancel{\text{mol }NH_3}}{3\ \cancel{\text{mol }H_2}} = 10.1666667\ \text{L }NH_3 \qquad \boxed{V_{NH_3} = 10.2\ \textbf{L } NH_3}$$

Example 5: A piece of aluminum metal is placed in an excess amount of sulfuric acid until all the metal is reacted. Calculate the mass of aluminum used if 5.76 L of hydrogen gas is evolved at STP.

$$2\ Al_{(s)} \qquad + \qquad 3\ H_2SO_{4\,(aq)} \quad \rightarrow \quad 3\ H_{2\,(g)} \qquad + \qquad Al_2(SO_4)_{3\,(aq)}$$
$$M = 26.98\ \text{g/mol} \qquad\qquad\qquad\qquad 5.76\ \text{L}$$
$$?\ \text{g} \qquad\qquad\qquad\qquad\qquad STP = 22.4\ \text{L/mol}$$

$$m_{Al} = 5.76\ \cancel{\text{L }H_2} \times \frac{1\ \cancel{\text{mol }H_2}}{22.4\ \cancel{\text{L }H_2}} \times \frac{2\ \cancel{\text{mol }Al}}{3\ \cancel{\text{mol }H_2}} \times \frac{26.98\ \text{g }Al}{1\ \cancel{\text{mol }Al}} = 4.625142857\ \text{g}$$

$$\boxed{m_{Al} = 4.63\ \textbf{g } Al}$$

Example 6: 35.24 kg of octane ($C_8H_{18\,(l)}$) is burned under excess oxygen. Determine the volume of carbon dioxide gas produced at STP.

$$2\ C_8H_{18\,(l)} \quad + \quad 25\ O_{2\,(g)} \quad \rightarrow \quad 16\ CO_{2\,(g)} \quad + \quad 18\ H_2O_{\,(g)}$$

35.24 kg   STP = 22.4 L/mol

M = 114.26 g/mol   ? L

$$V_{CO_2} = 35.24\ \text{kg}\ \cancel{C_8H_{18}} \times \frac{1\ \cancel{\text{mol}\ C_8H_{18}}}{114.26\ \text{g}\ \cancel{C_8H_{18}}} \times \frac{16\ \cancel{\text{mol}\ CO_2}}{2\ \cancel{\text{mol}\ C_8H_{18}}} \times \frac{22.4\ \text{L}\ CO_2}{1\ \cancel{\text{mol}\ CO_2}}$$

$$\boxed{V_{CO_2} = 55.3\ \text{kL}\ CO_2}$$
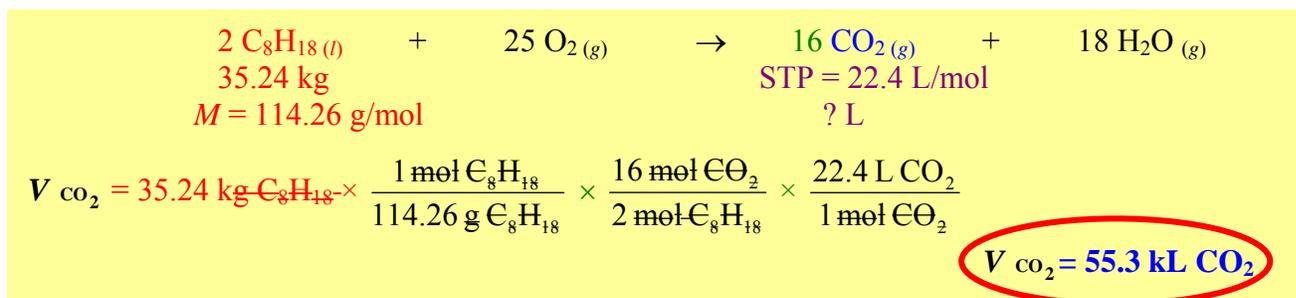
9.3: Calculations Involving Limiting Reagents

Excess: - the reactant with more than enough amount for the reaction.

Limiting Reagent: - the reactant with the smaller amount (after taken account of the mole ratio) for the reaction.

Note: A limiting reagent question will always have enough information to find the moles of both reactants.

Steps to deal with Limiting Reagent Problems:

1. Assume one of the reactants is the limiting reagent and determine its mole amount.

2. Determine the mole amount of the other reactant.

3. Use the mole amount of the assumed limiting reagent and the mole ratio; calculate the mole amount of the other reactant actually needed.

4. If the mole amount of the other reactant is smaller than what is needed, then our assumption was wrong. The other reactant is the limiting reagent.

5. If the mole amount of the other reactant is bigger than what is needed, then our assumption was correct. It means that the other reactant is the excess.

Example 1: A loaf of banana bread recipe calls for 3 eggs, 2 cups of flour, 1 cup of sugar and 3 bananas. How many loaves of banana bread can we make if we have 20 eggs, 11 cups of flour, 7 cups of sugar and 20 bananas? What is the limiting ingredient?

3 Eggs   +   2 Flour   +   1 Sugar   +   3 Bananas   →   1 Banana Bread

20 Eggs   11 Flour   7 Sugar   20 Bananas   ?

$$\frac{20\ \cancel{\text{Eggs}}}{3\ \cancel{\text{Eggs}}/\text{loave}} \qquad \frac{11\ \text{Flour}}{2\ \text{Flour}/\text{loave}} \qquad \frac{7\ \text{Sugar}}{1\ \text{Sugar}/\text{loave}} \qquad \frac{20\ \text{Bananas}}{3\ \text{Bananas}/\text{loave}}$$

= 6.67 loaves   **5.5 loaves**   = 7 loaves   = 6.67 loaves

(Minimum – Flour is the limiting ingredient)

Example 2: 5.00 g of phosphorus is reacted with 15.00 g of chlorine gas to produce phosphorus trichloride. Determine the mass of the product produced.

$$P_{4\,(s)} \quad + \quad 6\,Cl_{2\,(g)} \quad \rightarrow \quad 4\,PCl_{3\,(s)}$$

$$\begin{array}{ccc}
5.00\text{ g} & 15.00\text{ g} & ?\text{ g} \\
M = 123.88\text{ g/mol} & M = 70.90\text{ g/mol} & M = 137.32\text{ g/mol}
\end{array}$$
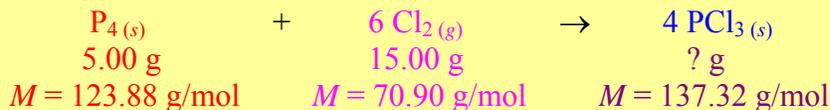
Since there is enough information to determine the moles of two reactants (quantities of both reactants are given), we need to find the mass of the product from each of these reactant before labelling which reactant is limiting.

① $5.00 \text{ g } P_4 \times \dfrac{1 \text{ mol } P_4}{123.88 \text{ g } P_4} \times \dfrac{4 \text{ mol } PCl_3}{1 \text{ mol } P_4} \times \dfrac{137.32 \text{ g } PCl_3}{1 \text{ mol } PCl_3} = 22.2 \text{ g } PCl_3$

② $15.00 \text{ g } Cl_2 \times \dfrac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \dfrac{4 \text{ mol } PCl_3}{6 \text{ mol } Cl_2} \times \dfrac{137.32 \text{ g } PCl_3}{1 \text{ mol } PCl_3} = \boxed{19.4 \text{ g } PCl_3}$ (lesser product mass)

Since Cl_2 gives a smaller calculated product mass, **Cl_2 is the limiting reactant**; **P_4 is the excess reactant.**

Assignment

9.1 pg. 240-241 #3, 4, 6 to 8

9.2 pg. 244 #9b, 10; pg. 245 #11, 12; pg. 248 #13, 14; pg. 249 #15, 16; pg. 250 #17 to 22

9.3 pg. 254 #23, 24; pg. 255 #25; pg. 256 #26; pg. 258 #27, 28; pg. 259 #29 to 32

Chapter 9 Review: pg. 262 #34 to 41, 43 to 47

Chapter 11: Thermochemistry – Heat and Chemical Change

11.1: The Flow of Energy–Heat

Thermochemistry: - the branch of chemistry that studies the amount of heat involved during a chemical change or a physical process.

Energy (E): - the ability to do **work** and/or produce **heat**.

Heat (q): - the transfer of energy between two objects (internal versus surroundings) due to the difference in temperature.

Work (w): - when force is applied over a displacement in the same direction ($w = F \times \Delta d$).
       - when a constant pressure is applied over a change in volume ($w = -P \, \Delta V$).

Temperature: - the **average kinetic energy** of all molecules / atoms in a substance.

Heat Capacity (Heat Content): - the **total amount of energy** of all molecules / atoms in a substance.

Chemical Potential Energy: - the energy that is "**stored**" within molecules / atoms.

Universe: - consists of both the **system** and the **surrounding**.

System: - a part of the entire universe as defined by the problem.

Surrounding: - the part of the universe outside the defined system.

First Law of Thermodynamics: - states that **energy cannot be created or destroyed**. It can only be converted from one form to another. Therefore, the total energy in the universe is constant.
       - also known as the **Law of Conservation of Energy** ($\Sigma E_{initial} = \Sigma E_{final}$).

Exothermic Process ($\Delta E < 0$): - when **energy flows "out" of the system** into the surrounding.
       (Surrounding gets Warmer.)

Potential Energy Diagram for Exothermic Process
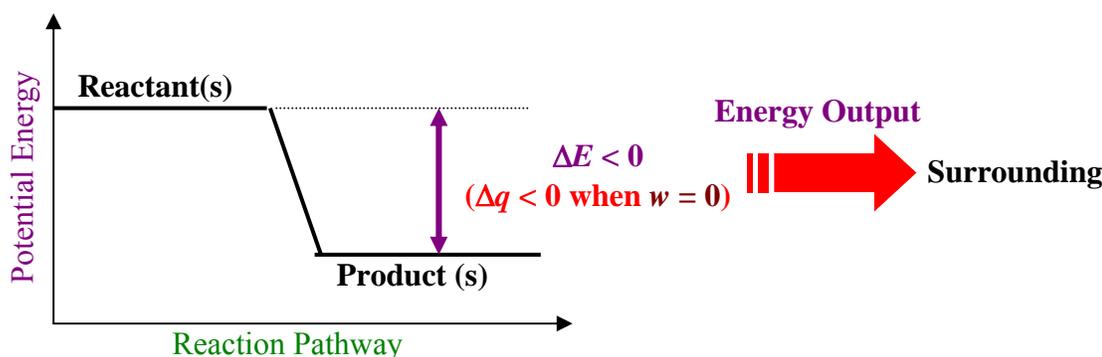
Endothermic Process ($\Delta E > 0$): - when **energy flows *"into"* the system** from the surrounding.
                                    (Surrounding gets Colder.)

Potential Energy Diagram for Endothermic Process
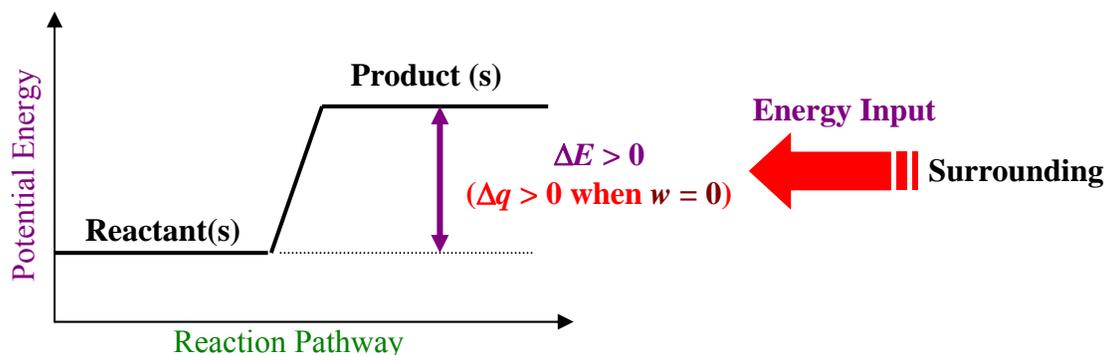
Physical Change: - when the temperature of a substance is raised or lowered, no new substance is formed.
            Only the **physical kinetic energy** is responsible for the heat change as its *molecules*
            are moving or vibrating faster or slower.
        - when a substance is changing phase, no new substance is formed. Only the **physical**
            potential energy is responsible for the heat change as its *intermolecular bonds are*
            being formed, broken or loosen.

Heat Units: - the units used to measure heat or energy.

a. Joules: - the metric unit to measure heat or energy named after the English physicist James Precott Joule.

b. Calories: - the old imperial unit to measure heat or energy. (1 cal = 4.19 J)

Specific Heat Capacity: - the amount of heat needed to raise one gram of substance by one degree Celsius.
                    - the higher the specific heat capacity, the more the substance can "hold" the heat.
                    - units are in J/(g • °C) or kJ/(kg • °C)

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Physical Kinetic Change

$$q = mC\Delta T$$

q = Change in Heat (J or kJ) m = mass (g or kg)
C = Specific Heat Capacity [J/(g • °C) or kJ/(kg • °C)]
ΔT = Change in Temperature (in °C)

</div>

Heat Capacity: - the amount of heat needed to raise a preset mass of substance by one degree Celsius.
            - commonly used when there are two different substances (like water and its container) are used.
            - it is basically the combined quantities of (***mC***) in the formula above.
            - units are in J/°C or kJ/°C

Specific Heat Capacity of Some Common Substances (at 1.00 atm and 298.15 K)

Substance	Specific Heat Capacity J/(g • °C) or kJ/(kg • °C)	Substance	Specific Heat Capacity J/(g • °C) or kJ/(kg • °C)
Ice $H_2O_{(s)}$	2.00	Aluminum $Al_{(s)}$	0.897
Water $H_2O_{(l)}$	4.19	Carbon (graphite) $C_{(s)}$	0.709
Steam $H_2O_{(g)}$	2.02	Copper $Cu_{(s)}$	0.385
Ammonia $NH_{3\,(g)}$	2.06	Iron $Fe_{(s)}$	0.449
Methanol $CH_3OH_{(l)}$	2.53	Silver $Ag_{(s)}$	0.235
Ethanol $C_2H_5OH_{(l)}$	2.44	Gold $Au_{(s)}$	0.129

Example 1: How much energy is needed to heat 100.0 g of water from 20.0°C to 80.0°C?

Since this question involves temperature (kinetic) change only, we need to use $q = mC\Delta T$.

$C = 4.19$ J/(g • °C)
$m = 100.0$ g H_2O
$\Delta T = 80.0°C - 20.0°C = 60.0°C$
$q = ?$

$q = mC\Delta T$
$q = (100.0 \text{ g})(4.19 \text{ J/(g • °C)})(60.0°C) = 25140$ J

$q = 2.51 \times 10^4$ J $= 25.1$ kJ

Example 2: A 20.0 g of an unknown element gave off 423. J of energy as it cooled from 100.0°C to 10.0°C. Identify this unknown element by determining its specific heat capacity.

$q = -423.$ J (q is negative because heat is given off)
$m = 20.0$ g
$\Delta T = 10.0°C - 100.0°C$
$\Delta T = -90.0°C$ ($\Delta T = T_{final} - T_{initial}$)
$C = ?$

$q = mC\Delta T$ $\dfrac{q}{m\Delta T} = C$

$C = \dfrac{(423. \text{ J})}{(20.0 \text{ g})(90.0°C)}$ $C = 0.235$ J/(g • °C)

From the reference table above, the unknown substance is silver.

Example 3: Calculate the final temperature of a 10.0 g piece of copper wire at 15.0°C when 3.27 kJ of heat is added.

$q = 3.27$ kJ
$m = 10.0$ g $= 0.0100$ kg of Cu
$C = 0.385$ kJ/(kg • °C)
$T_i = 15.0°C$

$\Delta T = T_f - 15.0°C = ?$

$q = mC\Delta T$ $\dfrac{q}{mC} = \Delta T$

$\Delta T = \dfrac{(3.27 \text{ kJ})}{(0.0100 \text{ kg})(0.385 \text{ kJ/(kg • °C)})}$ $\Delta T = 849.°C$

$\Delta T = T_f - 15.0°C$
$849.°C = T_f - 15.0°C$
$T_f = 849.°C + 15.0°C$ $T_f = 864.°C$

Example 4: Determine the energy involved when a container holding a preset amount of water has cooled from 85.0°C to 10.0°C and has a specific heat of 3.46 kJ/°C.

$(mC) = 3.46$ kJ/°C
$\Delta T = 10.0°C - 85.0°C = -75.0°C$
$q = ?$

$q = (mC)\Delta T$
$q = (3.46 \text{ kJ/°C})(-75.0°C) = -259.5$ kJ
(Negative because energy is given off)

$q = -260.$ kJ

Assignment
11.1 pg. 299 #1 to 10

11.3: Heat in Changes of State

Enthalpy (**H**): - the amount of internal energy (assuming there is no work) at a specific pressure and volume.

Energy involved in Physical Change (Temperature and / or Phase Change):

Heating Curve: - a graph of temeparture versus time as a substance is heated from a solid phase to a gaseous phase.
- when a substance is undergoing a **phase change**, its **temperature remains constant (the plateau on the heating curve)** until all molecules aquire enough energy to overcome the intermoelcular forces necessary. This is commonly referred to as the **potential change** of a subsatnce.
- when a substance is undergoing a **temperature change within a particular phase**, it is referred to as the **kinetic change** (because temeperature is defined as the average kinetic energy of a substance).

Heating Curve of Water
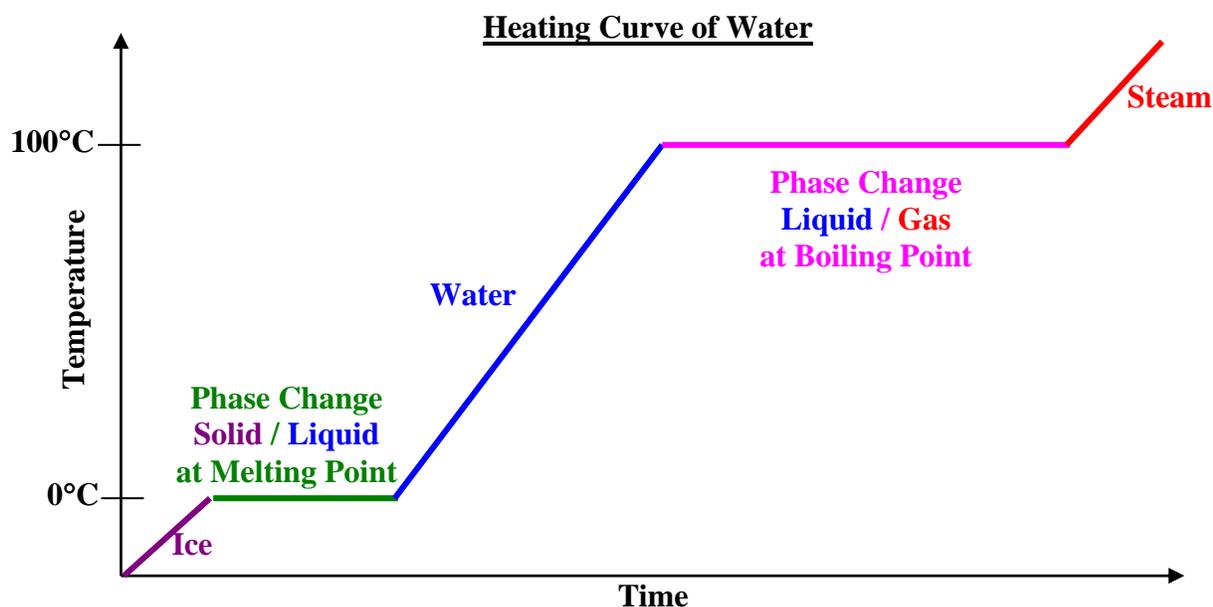
Molar Heat of Fusion (ΔH_{fus}): - the amount of heat needed to melt one mole of substance from solid to liquid at its melting point (in kJ/mol).

Molar Heat of Solidification (ΔH_{solid}): - the amount of heat needed to freeze one mole of substance from liquid to solid at its melting point (in kJ/mol).
- it is the same as ΔH_{fus} except the value of the rate is negative (heat is given off).

Molar Heat of Vaporization (ΔH_{vap}): - the amount of heat needed to evaporate one mole of substance from liquid to gas at its boiling point (in kJ/mol).

Molar Heat of Condensation ($\Delta H_{condense}$): - the amount of heat needed to condense one mole of substance from gas to liquid at its boiling point (in kJ/mol).
- it is the same as ΔH_{vap} except the value of the rate is negative (heat is given off).

Molar Heat of Solution (ΔH_{soln}): - the amount of heat needed to dissolve one mole of solid substance into water under standard condition of 25°C and at 1 atm (in kJ/mol).

Physical Potential Change

$$\Delta H = n\Delta H_{fus} \qquad \Delta H = n\Delta H_{solid} \qquad \Delta H = n\Delta H_{vap} \qquad \Delta H = n\Delta H_{condense} \qquad \Delta H = n\Delta H_{soln}$$

$$\Delta H_{fus} = -\Delta H_{solid} \qquad\qquad \Delta H_{vap} = -\Delta H_{condense}$$

ΔH = Change in Enthalpy n = moles ΔH_{soln} = Molar Heat of Solution (kJ/mol)

ΔH_{fus} = Molar Heat of Fusion (kJ/mol) ΔH_{vap} = Molar Heat of Vaporization (kJ/mol)

ΔH_{solid} = Molar Heat of Solidification (kJ/mol) $\Delta H_{condense}$ = Molar Heat of Condensation (kJ/mol)

Physical Thermodynamic Properties of Some Common Substances (at 1.00 atm and 298.15 K)

Substance	Melting Point (°C)	Boiling Point (°C)	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)
Ice $H_2O_{(s)}$	0	----	6.01	----
Water $H_2O_{(l)}$	----	100	----	40.65
Steam $H_2O_{(g)}$	----	----	----	----
Ammonia $NH_{3\,(g)}$	−77.73	−33.34	5.66	23.33
Methanol $CH_3OH_{(l)}$	−98	64.6	3.22	35.21
Ethanol $C_2H_5OH_{(l)}$	−114.1	78.3	4.93	38.56
Aluminum $Al_{(s)}$	660	2519	10.79	294.
Carbon (graphite) $C_{(s)}$	3338	4489	117.	----
Copper $Cu_{(s)}$	1085	2562	12.93	300.4
Iron $Fe_{(s)}$	1538	2861	13.81	340.
Mercury $Hg_{(l)}$	−39	357	2.29	59.1

Example 1: What is the change in enthalpy when 36.04 g of water condenses from gas to liquid at 100°C?

Since this question involves phase change (condensation) only, we need to use $\Delta H = n\Delta H_{condense}$.

$\Delta H_{condense} = -40.65$ kJ/mol

$n = \dfrac{36.04 \text{ g}}{18.02 \text{ g/mol}} = 2.000$ mol H_2O

$\Delta H = ?$

$\Delta H = n\Delta H_{condense}$

$\Delta H = (2.000 \text{ mol})(-40.65 \text{ kJ/mol})$

$\boxed{\Delta H = -81.3 \text{ kJ}}$ **(Negative; Heat is given off)**

Example 2: What is the total energy needed to heat 18.02 g of water at 80.0°C to steam at 115.0°C?

For this question, we have two kinetic changes (water and steam) and one phase change (vaporization).

$m = 18.02$ g $H_2O = 0.01802$ kg H_2O

$n = \dfrac{18.02 \text{ g}}{18.02 \text{ g/mol}}$

$n = 1.000$ mol H_2O

$C_{water} = 4.19$ kJ/(kg • °C)

$\Delta T_{water} = 100.0°C - 80.0°C = 20.0°C$

$\Delta H_{vap} = 40.65$ kJ/mol

$C_{steam} = 2.02$ kJ/(kg • °C)

$\Delta T_{steam} = 115.0°C - 100.0°C = 15.0°C$

$\Delta H = ?$

$\Delta H = mC_w\Delta T_w + n\Delta H_{vap} + mC_s\Delta T_s$
(water)   (vaporization)   (steam)

$\Delta H = (0.01802 \text{ kg})(4.19 \text{ kJ/(kg • °C)})(20.0°C) +$
$\qquad (1.000 \text{ mol})(40.65 \text{ kJ/mol}) +$
$\qquad (0.01802 \text{ kg})(2.02 \text{ kJ/(kg • °C)})(15.0°C)$

$\Delta H = 1.510076 \text{ kJ} + 40.65 \text{ kJ} + 0.546006 \text{ kJ}$

$\boxed{\Delta H = 42.7 \text{ kJ}}$

Writing ΔH Notations with Chemical Equations / Physical Process:

a. Endothermic Reactions / Processes

> **Reactant(s) + Heat → Product(s)   OR   Reactant(s) → Product(s) ΔH = +____ kJ**

Example: Water is vaporized from its liquid state.

> $H_2O_{(l)}$ + 40.65 kJ → $H_2O_{(g)}$ or $H_2O_{(l)}$ → $H_2O_{(g)}$ ΔH = +40.65 kJ

b. Exothermic Reactions / Processes

> **Reactant(s) → Product(s)   + Heat   OR   Reactant(s) → Product(s) ΔH = −____ kJ**

Example: Methane undergoes combustion at constant pressure.

$CH_{4(g)}$ + 2 $O_{2(g)}$ → $CO_{2(g)}$ + 2 $H_2O_{(g)}$ + 803. kJ or $CH_{4(g)}$ + 2 $O_{2(g)}$ → $CO_{2(g)}$ + 2 $H_2O_{(g)}$ ΔH = −803. kJ

Example 3: A cold pack consists of 40.0 g of NH_4NO_3 is dissolved in water. How much energy is absorbed or released into its surrounding if the ΔH_{soln} is 25.7 kJ/mol?

ΔH_{soln} = +25.7 kJ/mol (ΔH_{soln} > 0; endothermic) $\Delta H = n\Delta H_{soln}$
(Heat is absorbed from the surrounding) ΔH = (0.499625281 ~~mol~~)(25.7 kJ/~~mol~~)

$n = \dfrac{40.0\,g}{80.06\,g/mol} = 0.499625281$ mol NH_4NO_3

ΔH = ?   **ΔH = 12.8 kJ absorbed**

Example 4: 12.9 kJ of heat is released when $CaCl_2$ is dissolved in water, find the mass of $CaCl_2$ dissolved if the molar heat of solution of $CaCl_2$ is −82.8 kJ/mol.

ΔH_{soln} = −82.8 kJ/mol (ΔH_{soln} < 0; exothermic) $\Delta H = n\Delta H_{soln}$ $n = \dfrac{\Delta H}{\Delta H_{soln}}$
(Heat is released into the surrounding)
ΔH = −12.9 kJ
M = 110.98 g/mol $CaCl_2$ $n = \dfrac{-12.9\,kJ}{-82.8\,kJ/mol} = 0.1557971014$ mol

n = ? m = ? $m = nM$ = (0.1557971014 mol)(110.98 g/mol)

                    m = 17.3 g

Assignment
11.3 pg. 311 #22 and 23; pg. 313 #24 to 29

11.2: Measuring and Expressing Heat Changes

Molar Heat of Reaction (ΔH_{rxn}): - the amount of heat change **per mole of reactant used _or_ product** formed in a chemical reaction
- the molar heat of a combustion reaction is referred to as the **Molar Heat of Combustion (ΔH_{comb})**

Change in Enthalpy in a Chemical Reaction	**Chemical Potential Change**
$\Delta H = q = H_{products} - H_{reactants}$	$\Delta H = n\Delta H_{rxn}$
$\Delta H > 0$ **Endothermic Reaction** $\Delta H < 0$ **Exothermic Reaction**	ΔH = **Change in Enthalpy** n = **moles** ΔH_{rxn} = **Molar Enthalpy of Reaction (kJ/mol)**

Potential Energy Diagram for Endothermic Chemical Reactions / Physical Processes

Potential Energy

Product (s)

$\Delta H > 0$
($H_{products} > H_{reactants}$)

Reactant(s)

Reaction Pathway
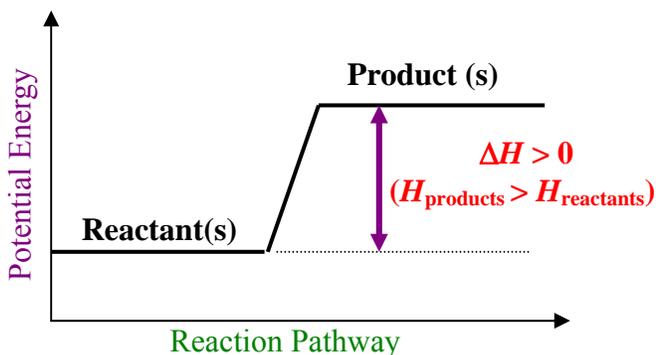
Potential Energy Diagram for Exothermic Chemical Reactions / Physical Processes

Potential Energy

Reactant(s)

$\Delta H < 0$
($H_{products} < H_{reactants}$)

Product (s)

Reaction Pathway
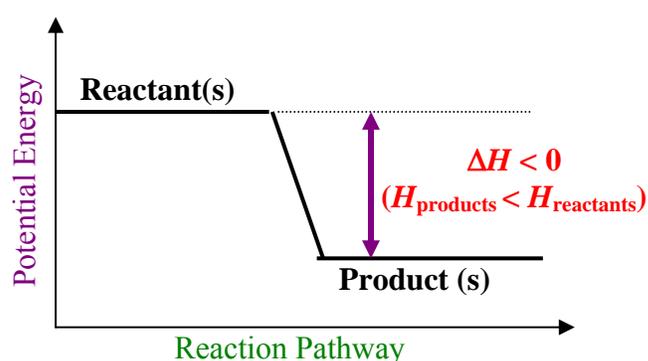
Example 1: Calculate the amount of energy generated when 500. g of octane is burned in a gasoline engine of a vehicle. ($\Delta H_{comb} = -5470.1$ kJ/mol)

$\Delta H_{comb} = -5470.1$ kJ/mol ($\Delta H_{comb} < 0$; exothermic)
(Heat is released into the surrounding)

$n = \dfrac{500.\,g}{114.26\,g/mol} = 4.375984597$ mol C_8H_{18}

$\Delta H = ?$

$\Delta H = n\Delta H_{comb}$
$\Delta H = (4.375984597 \; \text{mol})(-5470.1 \; \text{kJ/mol})$
$\Delta H = -23987$ kJ $= -2.40 \times 10^4$ kJ $= -24.0$ MJ

$\boxed{\Delta H = 2.40 \times 10^4 \text{ kJ} = 24.0 \text{ MJ released}}$

Example 2: $CH_2=CH_{2\,(g)} + H_2O_{(l)} \rightarrow C_2H_5OH_{(l)} + 44.2$ kJ. What is the mass of ethanol produced when 322. kJ is released from the reaction?

$\Delta H_{rxn} = -44.2$ kJ/mol ($\Delta H_{rxn} < 0$; exothermic)
(Heat is released into the surrounding)
$\Delta H = -322.$ kJ
$M = 46.08$ g/mol C_2H_5OH

$n = ?$ $m = ?$

$\Delta H = n\Delta H_{rxn}$ $n = \dfrac{\Delta H}{\Delta H_{rxn}}$

$n = \dfrac{-322.\,\text{kJ}}{-44.2\,\text{kJ/mol}} = 7.285067873$ mol

$m = nM = (7.285067873 \; \text{mol})(46.08 \; \text{g/mol})$

$\boxed{m = 336. \text{ g of } C_2H_5OH}$

Calorimetry: - uses the conservation of energy (**Heat Gained = Heat Lost**) to measure calories (old unit of heat: 1 cal = 4.18 J).
- physical calorimetry involves the mixing of two systems (one hotter than the other) to reach some final temperature.
- chemical calorimetry involves the heat change of a chemical reaction and the measurement of that change in the surrounding water's temperature.
- **the key to do these problems is to identify which system is gaining heat and which one is losing heat.**

A Simple Calorimeter   **Schematic of a Bomb Calorimeter**
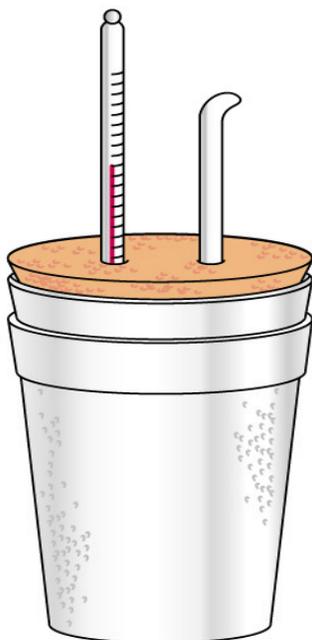
Motorized stirrer
Electrical leads for igniting sample
Thermometer
Insulated container
O_2 inlet
Bomb (reaction chamber)
Fine wire in contact with sample
Cup holding sample
Water
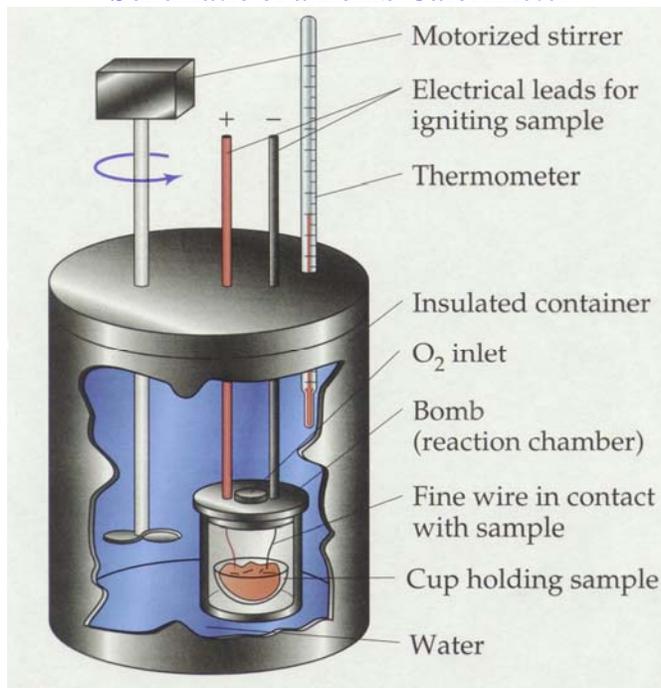
Physical Calorimetry
Heat Lost   =   Heat Gained
(Physical Change)   (water, kinetic)
$n\Delta H_{fus} = mC\Delta T$
or
$n\Delta H_{vap} = mC\Delta T$
or
$m_1 C_1 \Delta T_1 = m_w C_w \Delta T_w$

Chemical Calorimetry
Heat Lost   =   Heat Gained
(Chemical Reaction)   (water, kinetic)
$n\Delta H_{rxn} = mC\Delta T$
or
$n\Delta H_{sol} = mC\Delta T$

(Check out Calorimetry Animation at http://web.umr.edu/~gbert/animation.html)

Example 3: Hot water at 90.0°C is poured into 100. g of cold water at 10.0°C. The final temperature of the mixture is 70.0°C. Determine the mass of the hot water.

$m_{\text{cold water}} = 100.$ g
$C = 4.19$ J/(g • °C)
$\Delta T_{\text{hot water}} = 90.0°C - 70.0°C = 20.0°C$
(in calorimetry, we ignore negative temperature change)
$\Delta T_{\text{cold water}} = 70.0°C - 10.0°C = 60.0°C$

$m_{\text{hot water}} = ?$

Heat Lost   =   Heat Gained
(hot water, kinetic)   (cold water, kinetic)
$$m_{\text{hw}}\cancel{C}\Delta T_{\text{hw}} = m_{\text{cw}}\cancel{C}\Delta T_{\text{cw}}$$
$$m_{\text{hw}} = \frac{m_{\text{cw}}\Delta T_{\text{cw}}}{\Delta T_{\text{hw}}}$$
$$m_{\text{hw}} = \frac{(100.\,\text{g})(60.0°C)}{(20.0°C)}$$

$\boxed{m_{\text{hw}} = 300.\text{ g}}$

Example 4: A 1500. g of iron metal at 330°C is dropped into a beaker of 1000. g of water at 25.0°C. What will be the final temperature?

$T_f = ?$

$m_{\text{Fe}} = 1500.$ g
$C_{\text{Fe}} = 0.449$ J/(g • °C)
$\Delta T_{\text{Fe}} = 330°C - T_f$

$m_{\text{water}} = 1000.$ g
$C_{\text{water}} = 4.19$ J/(g • °C)
$\Delta T_{\text{water}} = T_f - 25.0°C$

Heat Lost   =   Heat Gained
(iron, kinetic)   (water, kinetic)
$$m_{\text{Fe}}C_{\text{Fe}}\Delta T_{\text{Fe}} = m_{\text{w}}C_{\text{w}}\Delta T_{\text{w}}$$
$(1500.\text{ g})(0.449\text{ J/(g • °C)})(330°C - T_f) = (1000.\text{ g})(4.19\text{ J/(g • °C)})(T_f - 25.0°C)$
$222255 - 673.5T_f = 4190T_f - 104750$
$-673.5T_f - 4190T_f = -104750 - 222255$
$-4863.5T_f = -327005$
$$T_f = \frac{-327005}{-4863.5}$$

$\boxed{T_f = 67.2°C}$

Example 5: When 10.02 g of liquid heptane, C_7H_{16}, is burned in the reaction vessel of a calorimeter, 1.50 L of water around the vessel increased its temperature from 20.0°C to 85.0°C. Ignore the metallic material of the calorimeter,

    a. determine the experimental molar heat of combustion heptane.
    b. calculate the % error in the calorimetry experiment for the combustion of heptane
        ($\Delta H_{comb} = -4.82$ MJ/mol for heptane burned).
    c. explain why the experimental ΔH is different than its theoretical counterpart.

a. We use the conservation of heat to calculate experimental ΔH_{rxn}.

$$n = \frac{10.02\,\text{g}}{100.23\,\text{g/mol}}$$
$n = 0.0999700688$ mol C_7H_{16}

$m_{water} = 1.50$ kg (1 kg = 1 L of water)
$\Delta T = 85.0°C - 20.0°C = 65.0°C$
$C_{water} = 4.19$ kJ / (kg • °C)

$\Delta H_{\text{rxn}} = ?$

$$n\Delta H_{\text{rxn}} = m_{\text{w}}C_{\text{w}}\Delta T$$
$$\Delta H_{\text{rxn}} = \frac{m_{\text{w}}C_{\text{w}}\Delta T}{n} = \frac{(1.50\,\cancel{\text{kg}})(4.19\,\text{kJ/}(\cancel{\text{kg}} • \cancel{°C}))(65.0\,\cancel{°C})}{(0.0999700688\,\text{mol})}$$

$\Delta H_{\text{rxn}} = 4086.473129$ kJ/mol (released)

$\boxed{\textbf{Experimental } \Delta H_{\textbf{rxn}} = \textbf{–4.09 MJ/mol of } C_7H_{16} \textbf{ burned}}$

b. $\% \text{ error} = \left| \dfrac{\text{Theoretical} - \text{Experimental}}{\text{Theroretical}} \right| \times 100\%$

$\% \text{ error} = \left| \dfrac{(-4.82 \text{ MJ}) - (-4.09 \text{ MJ})}{(-4.82 \text{ MJ})} \right| \times 100\%$ **% error = 15.1%**

c. Some of the possible reasons why experimental ΔH_{comb} (−4.08 MJ) is different than the theoretical ΔH_{comb} (−4.82 MJ)

➤ **Some of the heat released by the reaction is absorbed by the metal calorimeter itself. Thus, the energy gained by the water is not an exact reflection of the energy lost by the combustion.**

➤ **The calorimeter is not a closed system. Heat might escape into the surrounding.**

➤ **Even if the system is closed, the buildup of gases from the reaction would increase pressure and volume. Hence, some of the energy produced from the reaction is used to do work by the system. Thereby, it lowers the amount of heat available to warm the water.**

Assignment
11.2 pg. 302 #11 and 12; pg. 304 #13 and 14; pg. 306 #15 to 19

11.4: Calculating Heat Changes

Hess's Law: - the **addition of ΔH** of a series of reactions to find the change in enthalpy of an overall net reaction.
- when adding reactions, **compared the reactants and products of the overall net reaction with the intermediate reactions given**. **Decide on the intermediate reactions that need to be reversed or multiply by a coefficient or both, such that when added, the intermediate products will cancel out perfectly yielding the overall net reaction.**
- if a particular **reaction** needs to be **reversed (flipped)**, the **sign of the ΔH** for that reaction will also **need to be reversed**.
- if a **coefficient** is used to **multiply** a particular **reaction**, the **ΔH** for that reaction will also **have to multiply** by that same **coefficient**.

**(Check out Hess's Law Animation at
http://intro.chem.okstate.edu/1314F00/Lecture/Chapter6/Hesslaw2.html)**

Example 1: Calculate ΔH for the reaction $N_{2 (g)} + 2 \, O_{2 (g)} \rightarrow 2 \, NO_{2 (g)}$, when the following reactions are given.

$$N_{2 (g)} + O_{2 (g)} \rightarrow 2 \, NO_{(g)} \qquad \Delta H = 180 \text{ kJ}$$
$$2 \, NO_{2 (g)} \rightarrow 2 \, NO_{(g)} + O_{2 (g)} \qquad \Delta H = 112 \text{ kJ}$$
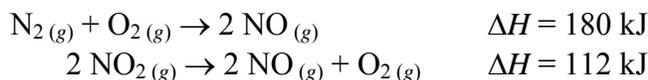
Note that $2 \, NO_2$ in the net reaction is on the product side, whereas $2 \, NO_2$ in the second reaction is on the reactant side. Hence, we need to reverse the second reaction and its sign of the ΔH.

$$N_{2 (g)} + O_{2 (g)} \rightarrow 2 \, NO_{(g)} \qquad \Delta H = \quad 180 \text{ kJ}$$
(Flipped) $2 \, NO_{(g)} + O_{2 (g)} \rightarrow 2 \, NO_{2 (g)} \qquad \Delta H = -112 \text{ kJ}$

$$N_{2 (g)} + 2 \, O_{2 (g)} \rightarrow 2 \, NO_{2 (g)} \qquad \Delta H = +68 \text{ kJ}$$
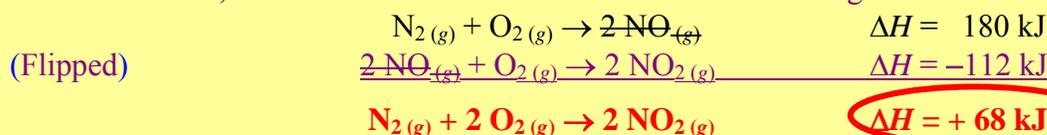
Example 2: Determine the ΔH for the reaction $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$, when the following reactions are given.

$$S_{(s)} + \frac{3}{2} O_{2(g)} \rightarrow SO_{3(g)} \qquad \Delta H = -395.2 \text{ kJ}$$

$$2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)} \qquad \Delta H = -198.2 \text{ kJ}$$
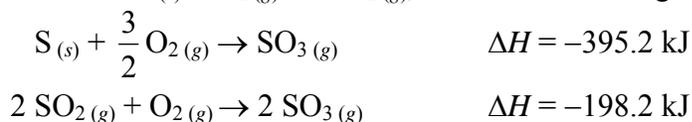
a. SO_2 in the net reaction is on the product side, whereas $2\ SO_2$ in the second reaction is on the reactant side. Hence, we need to reverse the second reaction and its sign of the ΔH.
b. There is only 1 SO_2 in the net reaction, whereas there are 2 SO_2 in the second reaction. Therefore the second reaction and its ΔH need to be multiplied by the coefficient ½.

$$S_{(s)} + \frac{3}{2} O_{2(g)} \rightarrow SO_{3(g)} \qquad \Delta H = \quad -395.2 \text{ kJ}$$

(Flipped and × ½)   ½ $(2\ SO_{3(g)} \rightarrow 2\ SO_{2(g)} + O_{2(g)})$ $\Delta H = ½(+198.2 \text{ kJ})$

$$S_{(s)} + \frac{3}{2} O_{2(g)} \rightarrow \cancel{SO_{3(g)}} \qquad \Delta H = -395.2 \text{ kJ}$$

$$\cancel{SO_{3(g)}} \rightarrow SO_{2(g)} + ½\ O_{2(g)} \quad \Delta H = \quad 99.1 \text{ kJ}$$

$$\mathbf{S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}} \qquad \boxed{\Delta H = -296.1 \text{ kJ}}$$
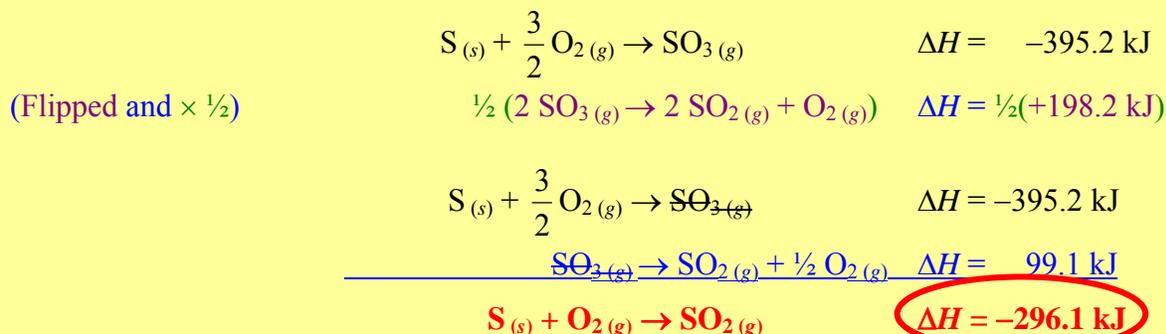
Example 3: Find the ΔH for the overall reaction of $2\ N_{2(g)} + 5\ O_{2(g)} \rightarrow 2\ N_2O_{5(g)}$, when the following reactions are given.

$$H_{2(g)} + ½\ O_{2(g)} \rightarrow H_2O_{(l)} \qquad \Delta H = -285.8 \text{ kJ}$$

$$N_2O_{5(g)} + H_2O_{(l)} \rightarrow 2\ HNO_{3(l)} \qquad \Delta H = -\ 76.6 \text{ kJ}$$

$$½\ N_{2(g)} + \frac{3}{2} O_{2(g)} + ½\ H_{2(g)} \rightarrow HNO_{3(l)} \qquad \Delta H = -174.1 \text{ kJ}$$
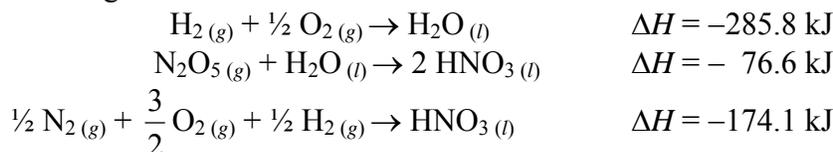
a. $2\ N_2O_5$ in the net reaction is on the product side, whereas N_2O_5 in the second reaction is on the reactant side. Hence, we need to reverse the second reaction and its sign of the ΔH.
b. There are $2\ N_2O_5$ in the net reaction, whereas there is only $1\ N_2O_5$ in the second reaction. Therefore the second reaction and its ΔH need to be multiplied by the coefficient 2.
c. There are $2\ N_2$ in the net reaction on the reactant side. Since $½\ N_2$ is on the reactant side of the third reaction, we need to multiply the third reaction and its ΔH by the coefficient 4.
d. In order for H_2O to cancel from the first and second reactions, we have to multiple the first reaction by 2 and flip. This is because H_2O in the second reaction has also flipped and has multiplied by 2.

(Flipped and × 2) $2\ (H_2O_{(l)} \rightarrow H_{2(g)} + ½\ O_{2(g)})$ $\Delta H = 2(+285.8 \text{ kJ})$
(Flipped and × 2) $2\ (2\ HNO_{3(l)} \rightarrow N_2O_{5(g)} + H_2O_{(l)})$ $\Delta H = 2(+\ 76.6 \text{ kJ})$
(× 4) $4\ (½\ N_{2(g)} + \frac{3}{2} O_{2(g)} + ½\ H_{2(g)} \rightarrow HNO_{3(l)})$ $\Delta H = 4(-174.1 \text{ kJ})$

$$\cancel{2\ H_2O_{(l)}} \rightarrow \cancel{2\ H_{2(g)}} + O_{2(g)} \qquad \Delta H = +571.6 \text{ kJ}$$

$$\cancel{4\ HNO_{3(l)}} \rightarrow 2\ N_2O_{5(g)} + \cancel{2\ H_2O_{(l)}} \qquad \Delta H = +153.2 \text{ kJ}$$

$$2\ N_{2(g)} + 6\ O_{2(g)} + \cancel{2\ H_{2(g)}} \rightarrow \cancel{4\ HNO_{3(l)}} \qquad \Delta H = -696.4 \text{ kJ}$$

$$\mathbf{2\ N_{2(g)} + 5\ O_{2(g)} \rightarrow 2\ N_2O_{5(g)}} \qquad \boxed{\Delta H = +\ 28.4 \text{ kJ}}$$
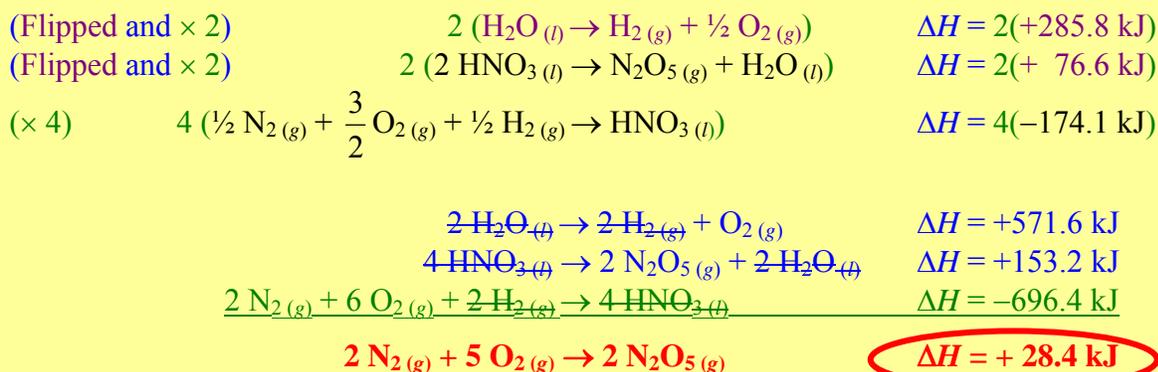
Molar Heat of Formation (ΔH_f): - the amount of heat required / given off to make 1 mole of compound from its elemental components.
- the **Molar Heat of Formation of ALL ELEMENTS is 0 kJ**.
- the **state of the compound affects the magnitude of H_f**.
  ($H_2O_{(g)}$ has $H_f = -241.8$ kJ/mol; $H_2O_{(l)}$ has $H_f = -285.8$ kJ/mol)
- **Standard Molar Heat of Formation ($H°_f$)** is the heat of formation under standard conditions (1 atm and 25°C).

(See Table 11.6 on pg. 316 in the Addison-Wesley 5th ed. Chemistry textbook for a list of $H°_f$)

Enthalpy of Formation (Chemical)

$$\Delta H = n\Delta H_f$$

ΔH = Change in Enthalpy   **n = moles**
ΔH_f = Molar Heat of Formation (kJ/mol)

Example 4: Find the molar heat of formation of table salt given that $2\ Na_{(s)} + Cl_{2\,(g)} \rightarrow 2\ NaCl_{(s)} + 822$ kJ.

$\Delta H = -822$ kJ
$n = 2$ mol of NaCl

$\Delta H_f = ?$

$\Delta H = n\Delta H_f$
$\Delta H_f = \dfrac{\Delta H}{n} = \dfrac{-822\ \text{kJ}}{2\ \text{mol}}$

$\boxed{\Delta H_f = -411\ \text{kJ/mol}}$

Example 5: What is the amount of heat absorbed / released when 100. g of $CO_{2\,(g)}$ is produced from its elements (CO_2 has $\Delta H_f = -393.5$ kJ/mol)?

$n = \dfrac{100.\,\text{g}}{44.01\,\text{g/mol}}$
$n = 2.272210861$ mol CO_2
$\Delta H_f = -393.5$ kJ/mol

$\Delta H = ?$

$\Delta H = n\Delta H_f$
$\Delta H = (2.272210861\ \text{mol})(-393.5\ \text{kJ/mol})$

$\boxed{\Delta H = -894.\ \text{kJ} \qquad (894.\ \text{kJ is released})}$

Example 6: Iron (III) oxide, rust, is produced from its elements, iron and oxygen. What is the mass of rust produced when 1.20 MJ is released as iron is reacted with oxygen? ($\Delta H_f = -824.2$ kJ/mol for Fe_2O_3)

$\Delta H = 1.20\ \text{MJ} = -1.20 \times 10^3$ kJ
$\Delta H_f = -824.2$ kJ/mol
$M = 159.70$ g/mol Fe_2O_3

$n = ?$ $m = ?$

$\Delta H = n\Delta H_f$
$n = \dfrac{\Delta H}{\Delta H_f} = \dfrac{-1.20 \times 10^3\ \text{kJ}}{-824.2\ \text{kJ/mol}} = 1.455957292$ mol

$m = nM = (1.455957292\ \text{mol})(159.70\ \text{g/mol})$

$\boxed{m = 233.\ \text{g}}$

Example 7: Calculate the molar heat of formation of silver (I) oxide when 91.2 g of Ag_2O is produced from its elements and 12.2 kJ of heat is released from the process.

$$n = \frac{91.2\,g}{231.74\,g/mol}$$

$n = 0.3935444895$ mol Ag_2O

$\Delta H = -12.2$ kJ

$\Delta H_f = ?$

$\Delta H = n\Delta H_f$

$$\Delta H_f = \frac{\Delta H}{n} = \frac{-12.2\,kJ}{0.3935444895\,mol}$$

$\boxed{\Delta H_f = -31.0\ \text{kJ/mol}}$

Molar Enthalpy of Reaction: - the amount of heat involved when 1 mol of a particular product is produced or 1 mol of a particular reactant is consumed.
- it is equal to **the difference between of all enthalpies of products and all enthalpies of reactants**.
- if the reaction is a **combustion**, it is called the **molar heat of combustion**.

(*Note*: **In hydrocarbon combustion, assume all products are gaseous unless otherwise stated.**)

Theoretical Enthalpy of Reaction

$$\Delta H_{rxn} = \Sigma H_{products} - \Sigma H_{reactants}$$

ΔH_{rxn} = **Change in Enthalpy of Reaction**
$\Sigma H_{products}$ = **Sum of Heat of Products (from all $n\Delta H_f$ of products)**
$\Sigma H_{reactants}$ = **Sum of Heat of Reactants (from all $n\Delta H_f$ of reactants)**

Example 8: Propane is a clean burning fossil fuel that is widely used in outdoor barbecue.
   a. Calculate the ΔH of combustion of propane.
   b. Draw its potential energy diagram.
   c. How much energy will be absorbed or released when 15.0 g or propane is burned?

a. We have to first write out a balance equation for the combustion of propane.

$$C_3H_{8\,(g)} + 5\,O_{2\,(g)} \rightarrow 3\,CO_{2\,(g)} + 4\,H_2O_{(g)}$$

$\Delta H_f:$ -103.8 kJ/mol   0 kJ/mol -393.5 kJ/mol -241.8 kJ/mol
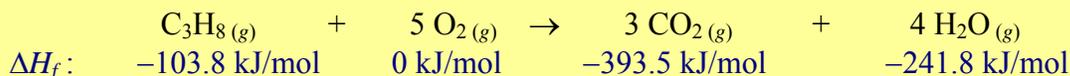
$\Delta H_{rxn} = \Sigma H_{products} - \Sigma H_{reactants}$
$\Delta H_{rxn} = [3\ \text{mol}\ (-393.5\ \text{kJ/mol}) + 4\ \text{mol}\ (-241.8\ \text{kJ/mol})] - [1\ \text{mol}\ (-103.8\ \text{kJ/mol}) + 5\ \text{mol}\ (0\ \text{kJ/mol})]$
$\Delta H_{rxn} = [-2147.7\ \text{kJ}] - [-103.8\ \text{kJ}]$

$\boxed{\Delta H_{rxn} = -2044.0\ \text{kJ/mol of } C_3H_8 \text{ burned}}$

b. Potential Energy Diagram
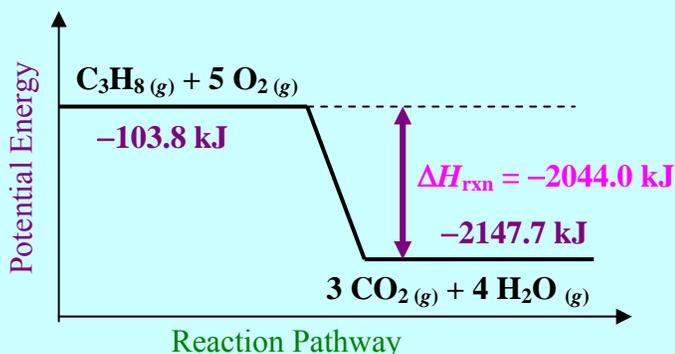
c. From part a., for every 1 mol of C_3H_8 burned, 2043.9 kJ is released.

$$n = \frac{15.0\,g}{44.11\,g/mol} = 0.3400589436 \text{ mol } C_3H_8$$

$\Delta H_{rxn} = -2044.0$ kJ/mol

$\Delta H = ?$

$\Delta H = n\Delta H_{rxn}$

$\Delta H = (0.3400589436 \text{ mol})(-2044.0 \text{ kJ/mol})$

$\boxed{\Delta H = -695. \text{ kJ}\quad(695. \text{ kJ is released})}$

Example 9: Determine the amount of heat released when 34.9 g of butane gas is burned ($\Delta H_f = -125.7$ kJ/mol for butane).

We have to first write out a balance equation for the combustion of butane.

$$2\,C_4H_{10\,(g)}\quad + \quad 13\,O_{2\,(g)}\quad \rightarrow \quad 8\,CO_{2\,(g)}\quad + \quad 10\,H_2O_{\,(g)}$$

(We have to divide all coefficients by 2 because we are calculating ΔH_{rxn} per mol of butane burned.)

$$C_4H_{10\,(g)}\quad + \quad \frac{13}{2}\,O_{2\,(g)}\quad \rightarrow \quad 4\,CO_{2\,(g)}\quad + \quad 5\,H_2O_{\,(g)}$$
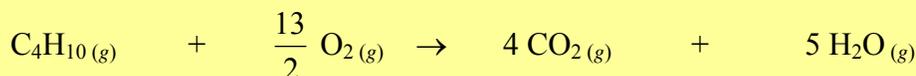
$\Delta H_f:\quad -125.7$ kJ/mol$\qquad$ 0 kJ/mol$\qquad$ -393.5 kJ/mol$\qquad$ -241.8 kJ/mol

$\Delta H_{rxn} = \Sigma H_{products} - \Sigma H_{reactants}$

$\Delta H_{rxn} = [4 \text{ mol }(-393.5 \text{ kJ/mol}) + 5 \text{ mol }(-241.8 \text{ kJ/mol})] - [1 \text{ mol }(-125.7 \text{ kJ/mol}) + \frac{13}{2}\text{ mol }(0 \text{ kJ/mol})]$

$\Delta H_{rxn} = [-2783.0 \text{ kJ}] - [-125.7 \text{ kJ}]$

$$\Delta H_{rxn} = -2657.0 \text{ kJ/mol of } C_4H_{10} \text{ burned}$$

$n = \dfrac{34.9\,g}{58.14\,g/mol} = 0.60027519878 \text{ mol } C_4H_{10}$

$\Delta H = n\Delta H_{rxn}$

$\Delta H = (0.60027519878 \text{ mol})(-2657.0 \text{ kJ/mol}) = -1594.93$ kJ

$\Delta H_{rxn} = -2657.0$ kJ/mol$\qquad \Delta H = ?$

$\boxed{\Delta H = -1.59 \text{ MJ}\quad(1.59 \text{ MJ is released})}$

Specific Heat in Calorimetry: - sometimes manufacturer of calorimeter has a preset volume of water to be filled for the calorimetry experiment. In that case, the calorimeter will come with its own specific heat value (in J/°C or kJ/°C), which is taken into account of the heat absorbed by the preset amount of water use as well as the calorimeter container itself.

Example 10: A calorimeter has a specific heat of 5.23 kJ/°C. What is the amount of heat absorbed when its temperature changes from 15.0°C to 50.0°C?

Since this question involves specific heat, we need to use $q = (mC)\Delta T$.

Specific Heat $= (mC) = 5.23$ kJ/°C

$\Delta T = 50.0°C - 15.0°C = 35.0°C$

$q = ?$

$q = (mC)\Delta T$

$q = (5.23 \text{ kJ/(°C)})(35.0°C) = 183.05$ kJ

$\boxed{q = 183. \text{ kJ}}$

Example 11: HCOOH $_{(l)}$ were completely burned into $CO_{2\ (g)}$ and $H_2O_{\ (l)}$ in a calorimeter. The following are the observations of the experiment.

Mass of HCOOH $_{(l)}$ burned	9.22 g
Initial Temperature of Calorimeter and Water	21.5°C
Final Temperature of Calorimeter and Water	37.3°C
Specific Heat of Calorimeter and Water	3.20 kJ/°C

    a. Determine the experimental molar heat of formation of HCOOH $_{(l)}$.
    b. If the theoretical ΔH_f for HCOOH $_{(l)}$ is −363. kJ/mol, calculate the % error of this experiment.

a. We use the conservation of heat to calculate the experimental ΔH_{rxn}.

$n = \dfrac{9.22\,\text{g}}{46.03\,\text{g/mol}}$

$n = 0.2003041495\,\text{mol HCOOH}$

$(mC)_{cal} = 3.20\,\text{kJ/°C}$
$\Delta T = 37.3°C - 21.5°C = 15.8°C$

$\Delta H_{rxn} = ?$

$n\Delta H_{rxn} = (mC)_{cal}\Delta T$

$\Delta H_{rxn} = \dfrac{(mC)_{cal}\,\Delta T}{n} = \dfrac{(3.20\,\text{kJ/°C})(15.8°C)}{(0.2003041495\,\text{mol})}$

$\Delta H_{rxn} = 252.4161388\,\text{kJ/mol (released)}$

Experimental ΔH_{rxn} = −252. kJ/mol of HCOOH burned

    Next, we use Hess's Law to find the ΔH_f of HCOOH.

$$\text{HCOOH}_{(l)} \quad + \quad \tfrac{1}{2}\,O_{2\ (g)} \quad \rightarrow \quad CO_{2\ (g)} \quad + \quad H_2O_{\ (l)}$$
$$\Delta H_f: \quad \textbf{?\ kJ/mol} \qquad 0\,\text{kJ/mol} \qquad -393.5\,\text{kJ/mol} \qquad -285.8\,\text{kJ/mol}$$

$\Delta H_{rxn} = \Sigma H_{products} - \Sigma H_{reactants}$
$-252.\,\text{kJ} = [1\ \text{mol}\ (-393.5\,\text{kJ/mol}) + 1\ \text{mol}\ (-285.8\,\text{kJ/mol})] - [1\ \text{mol}\ (\Delta H_f) + \tfrac{1}{2}\ \text{mol}\ (0\,\text{kJ/mol})]$
$-252.\,\text{kJ} = [-679.3\,\text{kJ}] - [\Delta H_f]$
    $\Delta H_f = -679.3\,\text{kJ} + 252.\,\text{kJ} = -427.3\,\text{kJ}$

Experimental HCOOH ΔH_f = −427. kJ/mol

b. % error $= \left|\dfrac{\text{Theoretical} - \text{Experimental}}{\text{Theroretical}}\right| \times 100\%$

% error $= \left|\dfrac{(-363.\,\text{kJ}) - (-427.\,\text{kJ})}{(-363.\,\text{kJ})}\right| \times 100\%$ **% error = 17.6%**

Assignment
11.4 pg. 317 #30 and 31; pg. 318 #32 to 35

11.5: Energy Productions and Environments

Greenhouse Effect: - the emission of greenhouse gases that traps more of the sun's radiant (heat) energy in the atmosphere.
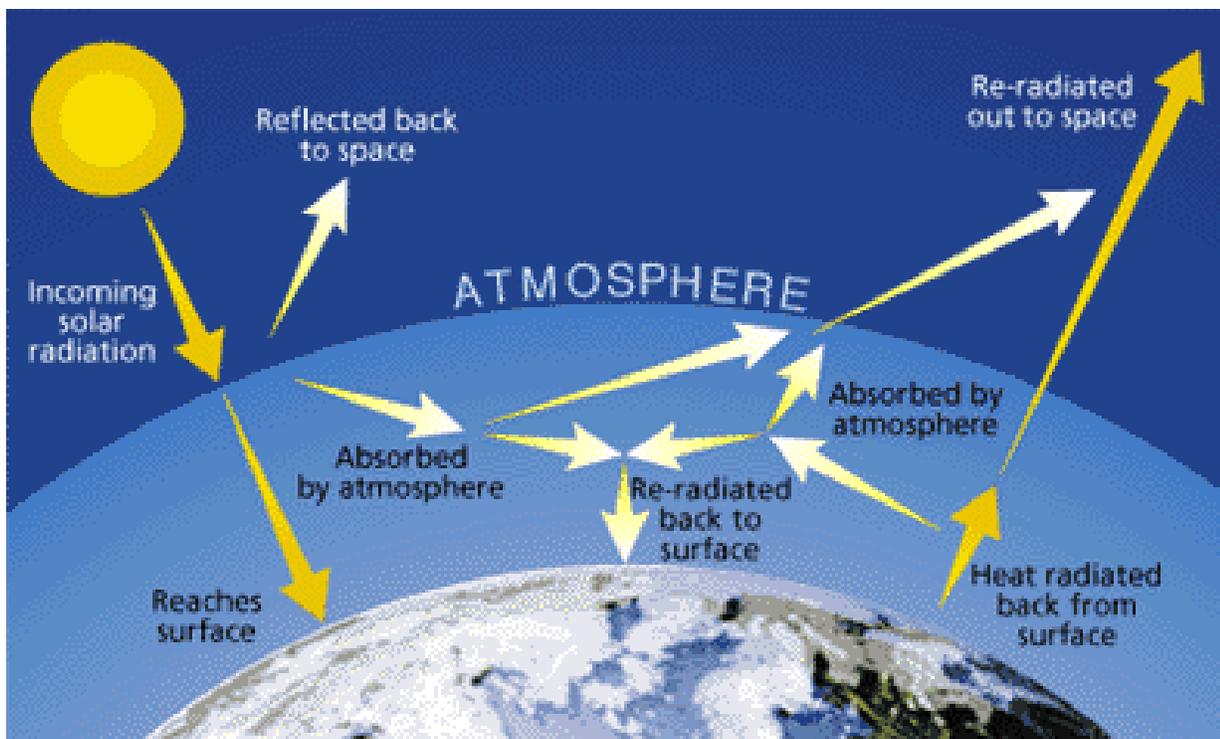
Greenhouses Gases: - man-made and naturally occurring gases that contribute to the Greenhouse Effect.

1. **Carbon dioxide (CO_2)** is released to the atmosphere when solid waste, fossil fuels (oil, natural gas, and coal), and wood and wood products are burned.

2. **Methane (CH_4)** is emitted during the production and transport of coal, natural gas, and oil. Methane emissions also result from the decomposition of organic wastes in municipal solid waste landfills, and the raising of livestock.

3. **Nitrous oxides (N_2O and NO_x)** are emitted during agricultural and industrial activities, as well as during combustion of solid wastes and fossil fuels.

4. **Hydrofluorocarbons (HFCs), Perfluorocarbons (PFCs), and Sulfur Hexafluoride (SF_6)** are very powerful greenhouse gases that do not occur naturally. They are generated in a variety of industrial processes.

    Each greenhouse gas differs in its ability to absorb heat in the atmosphere. HFCs and PFCs are the most heat-absorbent. Methane traps over 21 times more heat per molecule than carbon dioxide, and nitrous oxides absorb 270 times more heat per molecule than carbon dioxide. Often, estimates of greenhouse gas emissions are presented in units of millions of metric tons of carbon equivalents (MMTCE), which weigh each gas by its GWP value, or Global Warming Potential. (Information from US. EPA)

- Automobiles and Major Transportations account for 34% of CO_2 emissions globally (Power Plants contributes 33%; Major Industries and Home Heating contribute the remaining 33%).
- Presently 89% of all Energy Productions involve the burning of Fossil Fuels (Coal, Petroleum, Natural Gas and Biomass).
- Heat and Electricity generated from combustion of fossil fuel is at most 30% efficient.

  (Data from University of Michigan: http://www.umich.edu/~gs265/society/greenhouse.htm)

The Environmental Effect of Using Fossil Fuel: **(Greenhouse Effect)**

1. **Global Warming**: - the warming of global temperatures due to an increase amount of greenhouse gases in the atmosphere.

2. **Rise of Water Level**: - low-lying islands and coastal area are endangered as polar icecaps melt due to the rise of temperature as a result of global warming.

3. **Unpredicted and Erratic Climate**: - as the planet warms and polar icecaps melt, there is a change in salt-water concentrations in oceans as major ocean currents are modified. In some areas, there are increased precipitations, while other areas suffer persistent droughts.
   - high altitude glaciers are receding and thereby endangering the supply of fresh water to major tributaries (rivers).

4. **Deforestation**: - another cause of increasing CO_2 level in the atmosphere. As forests disappeared, fewer plants are available to absorb carbon dioxide using photosynthesis.
   - also causes mud and landslides, demineralization of the soil, lost of animal habitats and extinctions, destruction of entire ecosystems. Plants that may have important medicinal values can also be destroyed.

Alternate Energy Sources without the Emission of Greenhouse Gas

1. **Solar Energy**: - the most efficient energy source where energy from the sun is converted directly to electricity through the use of photovoltaic cells (solar panels) or heat using high efficient insulated glass and an effective water heating system.
   - technology exists but fairly expensive; requires many solar panels to generate adequate amount of electricity.

2. **Wind Power**: - the use of wind turbines to generate electricity.
   - very efficient and extremely economical, but location specific and not very reliable when there is no wind.
   - can disrupt migratory routes of birds (they get caught in the turbines), aesthetic problems for the landscapes.

3. **Geothermal Power**: - the use of underground steam (at volcanic or seismic regions) to generate electricity.
   - very efficient somewhat reliable, but location specific.
   - geothermal power is widely use in Iceland where it sits on the Atlantic ridge where there are lots of hot springs.

4. **Tidal Power**: - the use of tidal current to generate electricity.
- very efficient and somewhat reliable, but again location specific.
- tidal power involves establishing electric turbines at a narrow mouth of a channel where normal tides can cause bigger current and quick rise in water levels. It is being used in the Bay of Fundy at Nova Scotia, Canada and Kvalsund at the Arctic tip of Norway.
- tidal power can sometimes disrupt migratory routes of marine species.

5. **Hydroelectricity**: - the use of dam and reservoir to turn electric turbines as water falls from a higher level to the spillway (potential energy converted to kinetic energy to electricity).
- very efficient and no emission of greenhouse gas.
- location specific and very expensive to built. The reservoir flooding can destroy ecological habitats and force migrations of people from towns and villages (Aswan Dam in Egypt and the Three Gorges Dam in China displaced thousands of people and submerged ancient cities). The presence of the dam can disrupt aquatic migratory routes as well.
- dams have a limited life span (the collection of silt and mud at the bottom of the dam has to be clear periodically to maintain the structural integrity of the dam).
- dams can burst during earthquakes or poor maintenance. Flash flooding of towns along spillway is always a danger.

6. **Hydrogen Fuel**: - burning hydrogen to form water and generate heat and electricity.
- very efficient and zero pollution.
- hydrogen is very explosive and technologies are still needed for supplying and storing hydrogen safely in automobiles and homes.

Assignment
Chapter 11 Review: pg. 322−324 #36 to 48, 50 to 63, 69 to 75