1. Octane (C_8H_{18}) is a main component of gasoline used in cars.

   A) Write a balanced chemical equation for the combustion of octane with oxygen.
   $$\underline{2}\ C_8H_{18}\ (l) + \underline{25}\ O_2\ (g) \rightarrow \underline{16}\ CO_2\ (g) + \underline{18}\ H_2O\ (g)$$

   B) Using the following standard enthalpies of formation, determine the standard change in enthalpy (ΔH°_{rxn}) for the combustion reaction of octane.

   $$\Delta H^\circ_f[C_8H_{18}\ (l)] = -249.9\,\frac{kJ}{mol} \qquad \Delta H^\circ_f[CO_2\ (g)] = -393.5\,\frac{kJ}{mol} \qquad \Delta H^\circ_f[H_2O\ (g)] = -241.8\,\frac{kJ}{mol}$$

   $$\Delta H^\circ_{rxn} = n_{CO_2}\Delta H^\circ_f(CO_2) + n_{H_2O}\Delta H^\circ_f(H_2O) - n_{C_8H_{18}}\Delta H^\circ_f(C_8H_{18}) - n_{O_2}\Delta H^\circ_f(O_2)$$
   $$= 16\ mol\left(-393.5\,\frac{kJ}{mol}\right) + 18\ mol\left(-241.8\,\frac{kJ}{mol}\right) - 2\ mol\left(-249.9\,\frac{kJ}{mol}\right) - 25\ mol\left(0\,\frac{kJ}{mol}\right)$$
   $$\boldsymbol{\Delta H^\circ_{rxn} = -10148.6\ kJ}$$

   C) How much CO_2 gas is produced if 1.25 L of octane (density 0.703 g/mL) are combusted?
   $$1.25\ L\ C_8H_{18} \times \frac{1000\ mL}{1\ L} \times \frac{0.703\ g}{1\ mL} = 878._8\ g\ C_8H_{18}$$
   $$878._8\ g\ C_8H_{18} \times \frac{1\ mol}{114.22\ g} = 7.69_3\ mol\ C_8H_{18}$$
   $$7.69_3\ mol\ C_8H_{18} \times \frac{16\ mol\ CO_2}{2\ mol\ C_8H_{18}} \times \frac{44.01\ g\ CO_2}{1\ mol\ CO_2} = 2710\ g\ CO_2$$

2. A total of 1411 kJ of heat is applied to 500. mL of liquid water, originally 25.0 °C, to convert it all to water vapor. What is the final temperature of the gaseous water?

   $$c_P[H_2O(s)] = 37.1\,\frac{J}{mol \cdot °C} \qquad\qquad \Delta H_{fus}[H_2O] = 6.01\,\frac{kJ}{mol}$$

   $$c_P[H_2O(l)] = 75.3\,\frac{J}{mol \cdot °C} \qquad\qquad \Delta H_{vap}[H_2O] = 40.67\,\frac{kJ}{mol}$$

   $$c_P[H_2O(g)] = 33.6\,\frac{J}{mol \cdot °C}$$

   First, determine the amount of H₂O (l) using its density and molar mass:
   $$500.\ mL\ H_2O \times \frac{1\ g}{1\ mL} \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} = 27.7_5\ mol\ H_2O$$
   Second, determine how much heat is required to heat H₂O (l) from 25.0 °C to 100.0 °C:
   $$q_{liquid} = nc_P\Delta T = (27.7_5\ mol)\left(75.3\,\frac{J}{mol \cdot °C}\right)(100.0\ °C - 25.0\ °C) = 156700\ J = 156.7\ kJ$$
   Third, determine how much heat is required to vaporize H₂O (l) to H₂O (g):
   $$q_{vap} = n\Delta H_{vap} = (27.7_5\ mol)(40.67\ kJ) = 1128.6\ kJ$$
   Fourth, realize that we will heat the H₂O (g) after the vaporization process (q_{gas}), so:
   $$q_{total} = q_{liquid} + q_{vap} + q_{gas}$$
   $$q_{gas} = q_{total} - q_{liquid} - q_{vap} = 1411\ kJ - 156.7\ kJ - 1128.6\ kJ = 126\ kJ$$
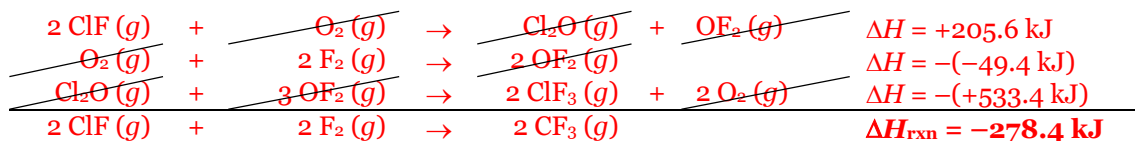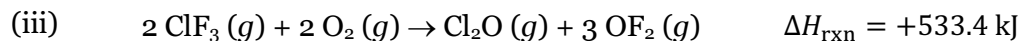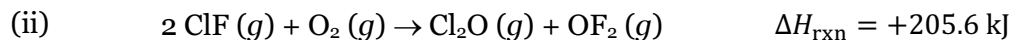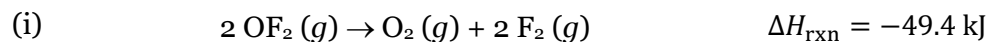   Lastly, determine the final temperature of the H₂O (g) from heating with 126 kJ heat:
   $$q_{gas} = nc_P\Delta T \rightarrow T_{final} - T_{initial} = \frac{q_{gas}}{nc_P}$$
   $$T_{final} = \frac{126\ kJ \times \frac{1000\ J}{1\ kJ}}{(27.7_5\ mol)\left(33.6\,\frac{J}{mol \cdot °C}\right)} + 100°C = \boldsymbol{230\ °C}$$

3. Consider the following reaction:

$$ClF\ (g) + F_2\ (g) \rightarrow ClF_3\ (g)$$
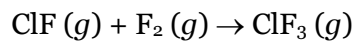
Calculate ΔH_{rxn} for the above reaction given the following reactions.

(i)	$2\ OF_2\ (g) \rightarrow O_2\ (g) + 2\ F_2\ (g)$	$\Delta H_{rxn} = -49.4\ kJ$
(ii)	$2\ ClF\ (g) + O_2\ (g) \rightarrow Cl_2O\ (g) + OF_2\ (g)$	$\Delta H_{rxn} = +205.6\ kJ$
(iii)	$2\ ClF_3\ (g) + 2\ O_2\ (g) \rightarrow Cl_2O\ (g) + 3\ OF_2\ (g)$	$\Delta H_{rxn} = +533.4\ kJ$

2 ClF (g)	+	~~O₂ (g)~~	→	~~Cl₂O (g)~~ + ~~OF₂ (g)~~		$\Delta H = +205.6\ kJ$
~~O₂ (g)~~	+	2 F₂ (g)	→	~~2 OF₂ (g)~~		$\Delta H = -(-49.4\ kJ)$
~~Cl₂O (g)~~	+	~~3 OF₂ (g)~~	→	2 ClF₃ (g) + ~~2 O₂ (g)~~		$\Delta H = -(+533.4\ kJ)$
2 ClF (g)	+	2 F₂ (g)	→	2 CF₃ (g)		$\Delta H_{rxn} = -278.4\ kJ$

But we need to divide the entire equation by 2, so $\Delta H_{rxn} = -139.2\ kJ$.