

Chapter 3: Thermochemistry

1. Which is a unit of energy?

- a. pascal
- b. newton
- * c. joule
- d. watt
- e. ampere

2. For a chemical reaction, where the internal energy is given the symbol E , the correct statement is:

- a. E_{final} signifies the internal energy of the reactants.
- b. E_{initial} signifies the internal energy of the products.
- * c. $\Delta E = E_{\text{products}} - E_{\text{reactants}}$
- d. ΔE is positive if energy is released to the surroundings.
- e. ΔE is positive if energy is released by the chemical reaction.

3. A 500.0 gram sample of aluminum is initially at 25.0 °C. It absorbs 32.60 kJ of heat from its surroundings. What is its final temperature, in °C?
(specific heat = 0.9930 J g⁻¹ °C⁻¹ for aluminum)

a. 40.4 °C b. 64.7 °C c. 65.7 °C d. 89.7 °C e. 90.7 °C	<p>Absorbs heat ($q > 0$) $q = m \times S \times (t_f - t_i)$ $+ 32.60 \times 10^3 \text{ J} = 500.0 \times 0.9930 \times (t_f - 25)$ $(t_f - 25) = 32600 / 500 \times 0.9930$ $(t_f - 25) = 65.66$ $t_f = 25 + 65.66 = 90.66 \text{ °C}$</p>
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4. A 113.25 gram sample of gold is initially at 100.0 °C. It loses 20.00 J of heat to its surroundings. What is its final temperature? (Specific heat of gold = 0.129 J g⁻¹ °C⁻¹)

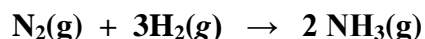
a. 98.6 °C b. -98.6 °C c. 94.6 °C d. -94.6 °C e. 96.6 °C	<p>Loses from system to surrounding ($q < 0$) $q = m \times S \times (t_f - t_i)$ $- 20.00 = 113.25 \times 0.129 \times (t_f - 100)$ $(t_f - 100) = -20.00 / 113.25 \times 0.129$ $(t_f - 100) = -1.4$ $t_f = 100 - 1.4 = 98.6 \text{ °C}$</p>
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5. During an **endothermic** chemical reaction,
- a system becomes warmer, and the chemical substances undergo an increase in potential energy.
 - a system becomes warmer, and the chemical substances undergo a decrease in potential energy.
 - * a system becomes **cooler**, and the chemical substances undergo an **increase in potential energy**.
 - a system becomes cooler, and the chemical substances undergo a decrease in potential energy.
 - a system becomes warmer, and additional heat is gained from the surroundings.

6. **Introducing** of 483 J of heat caused the system to **expand, doing** 320 J of work in the process against a constant pressure of 101 kPa (kilopascals). What is the value of ΔE for this process?

<ol style="list-style-type: none"> (483 + 320) joules (483 - 320) joules (320 - 483) joules 483 joules (-320 - 483) joules 	<p>Introducing heat = adding heat q > 0 Expand = system done work on surrounding w < 0</p> <p>$Q = + 483 \text{ J}, W = -320 \text{ J}$ $\Delta E = q + W$ $\Delta E = + 483 - 320$</p>
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7. A 92.38 kJ of heat are **given off** for each **mole of nitrogen gas** consumed. What is the the standard enthalpy of reaction in the equation below when 0.750 mol of hydrogen reacts?



<ol style="list-style-type: none"> +34.5 kJ -98.3 kJ +59.2 kJ -59.2 kJ -23.1 kJ 	<p>given off (exothermic) q < 0 $q = -92.38 \text{ kJ/mol}$</p> <p>$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$ 1 mol N₂(g) reacted with 3 mol H₂(g) ? 0.750 mol</p> <p>$n_{\text{N}_2} = 0.75/3 = 0.25 \text{ mol}$</p> <p>-92.38 kJ for 1 mol ? 0.25 mol $Q = -92.38 \text{ kJ} \times 0.25 = 23.1 \text{ kJ}$</p>
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8. The thermochemical equation which is associated with ΔH_f° , the standard enthalpy of formation for $\text{HCl}(g)$, is;

- a. $\text{H}(g) + \text{Cl}(g) \rightarrow \text{HCl}(g)$
- b. $\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2 \text{HCl}(g)$
- * c. $\frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{HCl}(g)$
- d. $\text{H}_2(g) + \text{Cl}_2(l) \rightarrow 2 \text{HCl}(g)$
- e. $\frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{Cl}_2(l) \rightarrow \text{HCl}(g)$

9. Consider the following thermochemical equation:

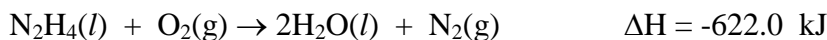
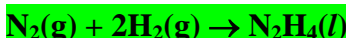


Calculate ΔH° for the reaction below:



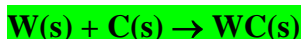
<ul style="list-style-type: none"> a. +334.5 kJ b. -146.19 kJ * c. +226.4 kJ d. -509.2 kJ e. +192.38 kJ 	Flip (reverse reaction) and double the first reaction to get the second reaction ΔH° (second reaction) = - 2 ΔH° (first reaction) = - 2 x (- 113.2 kJ) = + 226.4 kJ
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10. Determine the enthalpy change, ΔH , for the reaction,



<ul style="list-style-type: none"> a. -151.7 kJ b. -236.2 kJ c. +106.1 kJ * d. +50.2 kJ e. +567.4 kJ 	$2\text{H}_2\text{O}(l) + \text{N}_2(g) \rightarrow \text{N}_2\text{H}_4(l) + \text{O}_2(g) \quad \Delta H = + 622.0 \text{ kJ}$ $2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H = 2(-285.9 \text{ kJ})$ ----- $\text{N}_2(g) + 2\text{H}_2(g) \rightarrow \text{N}_2\text{H}_4(l) \quad \Delta H = 622.0 - 2(285.9) = +50.2 \text{ kJ}$
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11. Determine ΔH , for the reaction,



Given the following thermochemical equations:

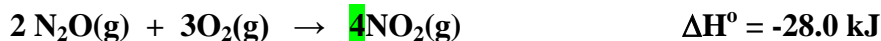


<ul style="list-style-type: none"> a. +33.3 kJ * b. -38.2 kJ c. +106.1 kJ d. -52.9 kJ e. +177.4 kJ 	$\text{W}(s) + \frac{3}{2} \text{O}_2(g) \rightarrow \text{WO}_3(s) \quad \Delta H = -1680.8 / 2 = -840.4 \text{ kJ}$ $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H = -393.5 \text{ kJ}$ $\text{WO}_3(s) + \text{CO}_2(g) \rightarrow \text{WC}(s) + \frac{5}{2} \text{O}_2(g) \quad \Delta H = 2391.4 / 2 = +1195.7 \text{ kJ}$ ----- $\Delta H = -840.4 \text{ kJ} - 393.5 \text{ kJ} + 1195.7 \text{ kJ} = -38.2 \text{ kJ}$
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12. The mathematical equation which expresses the first law of thermodynamics is;

- a. $\Delta H = \Delta E + p\Delta V$
- b. $\Delta H = \Delta E - p\Delta V$
- c. $\Delta H = q + w$
- * d. $\Delta E = q + w$
- e. $\Delta H = q + \Delta E$

13. Write the equation for the reaction when 1.00 mole of NO_2 is formed.



Divide the equation by 4 and $-28/4 = -7 \text{ kJ}$

- a. $\text{N}_2\text{O}(\text{g}) + 3 \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \quad \Delta H^\circ = -28.0 \text{ kJ}$
- b. $\text{N}_2\text{O}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \quad \Delta H^\circ = -28.0 \text{ kJ}$
- c. $2 \text{N}_2\text{O}(\text{g}) + 3 \text{O}_2(\text{g}) \rightarrow 4 \text{NO}_2(\text{g}) \quad \Delta H^\circ = -56.0 \text{ kJ}$
- d. $\frac{1}{2} \text{N}_2\text{O}(\text{g}) + \frac{3}{4} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \quad \Delta H^\circ = -7.00 \text{ kJ}$
- e. $\frac{1}{2} \text{N}_2\text{O}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \quad \Delta H^\circ = -14.0 \text{ kJ}$

14. Given the equation, $3\text{A} + 4\text{B} \rightarrow 4\text{C} + 7\text{D}$, and the following standard enthalpies of formation, ΔH_f° :

A: $+15.7 \text{ kJ mol}^{-1}$ B: $-86.4 \text{ kJ mol}^{-1}$ C: $-52.7 \text{ kJ mol}^{-1}$ D: $-71.6 \text{ kJ mol}^{-1}$

Calculate the standard enthalpy of reaction, in kJ, for the reaction shown.

a. -53.6 kJ	$\Delta H^\circ = 4 \Delta H_f^\circ (\text{C}) + 7 \Delta H_f^\circ (\text{D}) - 3 \Delta H_f^\circ (\text{A}) - 4 \Delta H_f^\circ (\text{B})$
b. -413.5 kJ	$\Delta H^\circ = 4 \times (-52.7) + 7(-71.6) - 3(+15.7) - 4(-86.4)$
c. -515.6 kJ	$\Delta H^\circ = -413.5 \text{ kJ}$
d. -853.6 kJ	
e. -908.4 kJ	

15. The standard enthalpy of combustion for xylene, $\text{C}_8\text{H}_{10}(\text{l})$, is $-3908 \text{ kJ mol}^{-1}$. Using ΔH_f° : $\text{H}_2\text{O}(\text{l}) = -285.9 \text{ kJ mol}^{-1}$; $\text{CO}_2(\text{g}) = -393.5 \text{ kJ mol}^{-1}$, to calculate ΔH_f° $\text{C}_8\text{H}_{10}(\text{l})$, in kJ mol^{-1} .



a. -669.5 kJ	$\Delta H^\circ(\text{comb.}) = 5 \Delta H_f^\circ (\text{H}_2\text{O}) + 8 \Delta H_f^\circ (\text{CO}_2) - \Delta H_f^\circ (\text{C}_8\text{H}_{10}) - 10.5 \Delta H_f^\circ (\text{O}_2)$
b. $+3228.6 \text{ kJ}$	$-3908 = 5 (-285.9) + 8 (-393.5) - \Delta H_f^\circ (\text{C}_8\text{H}_{10}) - 10.5 (0)$
c. -3228.6 kJ	$\Delta H_f^\circ (\text{C}_8\text{H}_{10}) = 3908 + 5 (-285.9) + 8 (-393.5) - 10.5 (0)$
d. $+4587.4 \text{ kJ}$	
e. $+8485.5 \text{ kJ}$	$\Delta H_f^\circ (\text{C}_8\text{H}_{10}) = -669.5 \text{ kJ}$

16. The standard enthalpy of reaction, $\Delta H^\circ_{\text{rxn}}$ for

$$\text{C}_2\text{H}_2(\text{g}) + 2 \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$$
 is $-311.5 \text{ kJ mol}^{-1}$. Determine the value of $\Delta E^\circ_{\text{rxn}}$ for this reaction.

* a. $-306.5 \text{ kJ mol}^{-1}$	$\Delta E^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - P \Delta V = \Delta H^\circ_{\text{rxn}} - \Delta n(\text{g}) RT$ at standard condition
b. $-309.0 \text{ kJ mol}^{-1}$	($T=298\text{K}$) and $\Delta n = \sum n(\text{g})_{\text{Products}} - \sum n(\text{g})_{\text{Reactants}} = 1-2-1 \text{ mol}$
c. $-314.0 \text{ kJ mol}^{-1}$	$\Delta E^\circ_{\text{rxn}} = -311.5 \text{ kJ mol}^{-1} - [(1-2-1) \times 8.314 \times 10^{-3} \text{ kJ/mol.K} \times 298\text{K}]$
d. $-316.46 \text{ kJ mol}^{-1}$	$\Delta E^\circ_{\text{rxn}} = -311.5 + (2 \times 8.314 \times 10^{-3} \times 298)$
e. $+4646 \text{ kJ mol}^{-1}$	$\Delta E^\circ_{\text{rxn}} = -306.5 \text{ kJ mol}^{-1}$

Extra problems

17.

How can energy be transferred to or from a system?

- A) Energy can only be transferred as potential energy being converted to kinetic energy.
- B) Energy can be transferred only as heat.
- C) Energy can be transferred only as work.
- D) Energy can be transferred as heat and/or work.

18.

Which of the following is NOT a state property?

- A) pressure
- B) temperature
- C) internal energy
- D) enthalpy
- E) work

State function such as ΔE , ΔH , ΔT , ΔP , ΔV , Δn which is depend on the initial and final value = $(X_f - X_i)$.

But not state function w, q

19.

Which of the following is the best example of an isolated system?

- A) water in a styrofoam coffee cup
- B) liquid in a beaker with a watch glass over it
- C) soda in an unopened soft drink can
- D) coffee in a closed thermos bottle

20.

How does a closed system differ from an open system?

- A) A closed system does not do any work on the surroundings.
- B) A closed system cannot exchange heat with the surroundings.
- C) Energy is conserved in a closed system, but not in an open system.
- D) A closed system cannot exchange matter with the surroundings.

21.

A chemical reaction where heat is transferred to the surroundings is a(n) _____ reaction.

EXOTHERMIC

22.

Which of the following is an endothermic process?

- A) jet fuel burning in a jet engine
- B) combustion of methane
- C) freezing of water
- D) vaporization of water

Breaking bonds is endothermic (vaporization mean change water from liquid to gas, breaking bonds $q > 0$)

23.

Which is true if $\Delta H = -95 \text{ J}$?

- A) Both the system and the surroundings are gaining 95 J.
- B) Both the system and the surroundings are losing 95 J.
- C) The system is gaining 95 J, while the surroundings are losing 95 J.
- D) The system is losing 95 J, while the surroundings are gaining 95 J.
- E) The system is losing -95 J, while the surroundings are gaining 95 J.

The reaction exothermic ($\Delta H < 0$) and the heat transferred from system to surrounding

24.

Which of the following signs on q and w represent a system that is doing work on the surroundings, as well as losing heat to the surroundings?

A) $+q, +w$

B) $-q, -w$

C) $+q, -w$

D) $-q, +w$

E) None of the above.

Losing from system ($q < 0$) and doing work ($w < 0$)

25.

How much work in joules is done when a piston expands from a volume of 13.27 liters to 76.55 liters against a pressure of 14.89 atm?

$$w = -P\Delta V$$

$$w = -(14.89 \text{ atm})(76.55 \text{ L} - 13.27 \text{ L})$$

$$w = -942.2 \text{ L}\cdot\text{atm}$$

$$-942.2 \text{ L}\cdot\text{atm} \times \frac{101.3 \text{ J}}{\text{L}\cdot\text{atm}} = -95450 \text{ J}$$

Expansion work done by the system on surrounding ($w < 0$), $\text{L}\cdot\text{atm} \times 101.325 = \text{J}$

26.

In a refrigeration system, the refrigerant gas absorbs 21.39 kJ of energy while expanding against a 0.278 atmosphere pressure from a volume of 0.0423 liters to a volume of 1.876 liters. What is the energy change of the gas?

$$\Delta E = \Delta H - P\Delta V = (21390 \text{ J}) - (0.278 \text{ atm})(1.876 \text{ L} - 0.0423 \text{ L}) \left(\frac{101.3 \text{ J}}{\text{L}\cdot\text{atm}} \right)$$

$$\Delta E = 21338 \text{ J} \Rightarrow 21.338 \text{ kJ}$$

System absorbed heat ($q > 0$), Expansion means work done by the system on surrounding ($w < 0$) $\text{L}\cdot\text{atm} \times 101.325 = \text{J}$

27.

How much work in joules is done on the system when a 1.15 atm external pressure causes a piston to decrease in volume from 6.55 liters to 3.16 liters?

$$w = -P\Delta V$$

$$w = -(1.15 \text{ atm})(3.16 \text{ L} - 6.55 \text{ L})$$

$$w = 3.90 \text{ L}\cdot\text{atm}$$

$$3.90 \text{ L}\cdot\text{atm} \times \frac{101.3 \text{ J}}{\text{L}\cdot\text{atm}} = 395 \text{ J}$$

Work done on the system $w > 0$, (decreasing volume)

$\text{L}\cdot\text{atm} \times 101.325 = \text{J}$

28.

What is the energy change of the system if a chemical reaction transfers 32.146 kJ of heat to the surroundings while it causes the expansion of a 1.465 liter vessel to 3.687 liters against a pressure of 3.64 atmospheres?

$$\Delta E = -32.965 \text{ kJ}$$

Heat transferred from system to surrounding $q < 0$, and expansion, work done by the system

$$w = -p\Delta V = -3.64(3.687-1.465) \times 101.325 \times 10^{-3} \text{ L}\cdot\text{atm} \times 101.325 \times 10^{-3} = \text{kJ}$$

$$q = -32.146 \text{ kJ}$$

$$\Delta E = Q + W = [-3.64(3.687-1.465) \times 101.325 \times 10^{-3}] + [-32.146] = -32.965 \text{ kJ}$$

29.

Lead, water, sulfur, and arsenic have specific heats of 0.128, 4.18, 0.706, and 0.329 J/g°C, respectively. Which of these would require the smallest amount of heat to increase its temperature by 10 °C (assume all samples have the same mass)?

LEAD \Rightarrow SMALLEST SPECIFIC HEAT

As specific heat decreases, less amount of heat will increase required reaching required temperature

30.

14.0 g of metal at 24.0 °C has 250 joules of heat added to it. The metal's specific heat is 0.105 J/g °C. What is its final temperature?

$$q = mC\Delta T \Rightarrow 250 \text{ J} = (14.0 \text{ g})(0.105 \frac{\text{J}}{\text{g}^\circ\text{C}})(T_f - 24.0^\circ\text{C})$$

$$T_f = 28.3^\circ\text{C}$$

Heat added $q > 0$, then the temperature will increase

31.

1219 joules of heat raise the temperature of 250 g of metal by 64 °C. What is the specific heat in J/g °C?

$$c = \frac{q}{m \Delta T} = \frac{(1219 \text{ J})}{(250 \text{ g})(64^\circ \text{C})}$$
$$c = 0.076 \frac{\text{J}}{\text{g}^\circ \text{C}}$$

Raise temperature, heat required, $q > 0$

32.

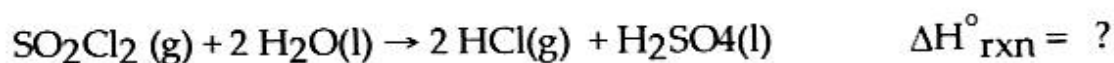
Two aqueous solutions at room temperature are mixed in a coffee cup calorimeter. The reaction causes the temperature of the resulting solution to fall below room temperature. Which of the following statements is TRUE?

- A) Energy is leaving the system during reaction.
- B) The products have a lower potential energy than the reactants.
- C) This type of experiment directly yields ΔE_{rxn} .
- D) The mixing is endothermic.
- E) The solution has special properties that enable it to violate the first and second law of thermodynamics.

Reaction takes heat from surrounding (water solution), and temperature of the solution will decrease

33.

Use the ΔH°_f information provided to calculate $\Delta H^\circ_{\text{rxn}}$ for the following:



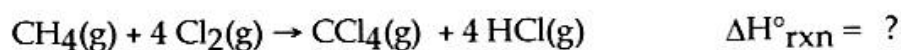
	ΔH°_f (kJ/mol)
$\text{SO}_2\text{Cl}_2(\text{g})$	-364
$\text{H}_2\text{O}(\text{l})$	-286
$\text{HCl}(\text{g})$	-92
$\text{H}_2\text{SO}_4(\text{l})$	-814

$$\Delta H^{\circ}_{\text{rxn}} = \left[(2 \text{ mol}) \left(-92 \frac{\text{kJ}}{\text{mol}} \right) + (1 \text{ mol}) \left(-814 \frac{\text{kJ}}{\text{mol}} \right) \right] - \left[(1 \text{ mol}) \left(-364 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol}) \left(-286 \frac{\text{kJ}}{\text{mol}} \right) \right] = -62 \text{ kJ}$$

$$\Delta H^{\circ} = 2 \Delta H^{\circ}_f(\text{HCl}) + 1 \Delta H^{\circ}_f(\text{H}_2\text{SO}_4) - 2 \Delta H^{\circ}_f(\text{H}_2\text{O}) - 1 \Delta H^{\circ}_f(\text{SO}_2\text{Cl}_2)$$

34.

Use the information provided to determine $\Delta H^{\circ}_{\text{rxn}}$ for the following reaction:



	ΔH°_f (kJ/mol)
CH ₄ (g)	-75
CCl ₄ (g)	-96
HCl(g)	-92

$$= \left[(1 \text{ mol}) \left(-96 \frac{\text{kJ}}{\text{mol}} \right) + (4 \text{ mol}) \left(-92 \frac{\text{kJ}}{\text{mol}} \right) \right] - \left[(1 \text{ mol}) \left(-75 \frac{\text{kJ}}{\text{mol}} \right) \right] = -389 \text{ kJ}$$

$$\Delta H^{\circ} = 4 \Delta H^{\circ}_f(\text{HCl}) + 1 \Delta H^{\circ}_f(\text{CCl}_4) - 4 \Delta H^{\circ}_f(\text{Cl}_2) - 1 \Delta H^{\circ}_f(\text{CH}_4)$$

$$\Delta H^{\circ}_f(\text{Cl}_2) = 0$$

35.

Use the information provided to determine $\Delta H^{\circ}_{\text{rxn}}$ for the following reaction:



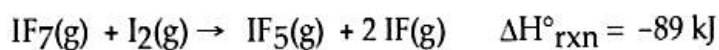
	ΔH°_f (kJ/mol)
Fe ₂ O ₃ (s)	-824
Fe ₃ O ₄ (s)	-1118
CO(g)	-111
CO ₂ (g)	-394

$$\Delta H^{\circ}_{\text{rxn}} = \left[(2 \text{ mol}) \left(-1118 \frac{\text{kJ}}{\text{mol}} \right) + (1 \text{ mol}) \left(-394 \frac{\text{kJ}}{\text{mol}} \right) \right] - \left[(3 \text{ mol}) \left(-824 \frac{\text{kJ}}{\text{mol}} \right) + (1 \text{ mol}) \left(-111 \frac{\text{kJ}}{\text{mol}} \right) \right] = -47 \text{ kJ}$$

$$\Delta H^\circ = 2 \Delta H_f^\circ (\text{Fe}_3\text{O}_4) + 1 \Delta H_f^\circ (\text{CO}_2) - 3 \Delta H_f^\circ (\text{Fe}_2\text{O}_3) - 1 \Delta H_f^\circ (\text{CO})$$

36.

Use the ΔH_f° and $\Delta H_{\text{rxn}}^\circ$ information provided to calculate ΔH_f° for IF:



$\text{IF}_7(\text{g})$	$\frac{\Delta H_f^\circ}{\text{kJ/mol}}$	$-89 \text{ kJ} = \left[(1 \text{ mol})(-840 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol})(\Delta H_f^\circ_{\text{IF}}) \right] - \left[(1 \text{ mol})(-941) \right]$
	-941	
$\text{IF}_5(\text{g})$	-840	

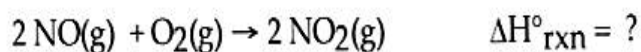
$$\Delta H_f^\circ_{\text{IF}} = -95 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta H^\circ = 2 \Delta H_f^\circ (\text{IF}) + 1 \Delta H_f^\circ (\text{IF}_3) - 1 \Delta H_f^\circ (\text{IF}_7) - 1 \Delta H_f^\circ (\text{I}_2)$$

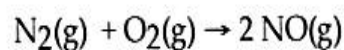
$$\Delta H_f^\circ (\text{I}_2) = 0$$

37.

Use the standard reaction enthalpies given below to determine $\Delta H^\circ_{\text{rxn}}$ for the following reaction:

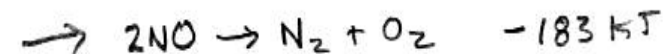


Given:

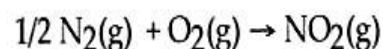


$$\Delta H^\circ_{\text{rxn}} = 183 \text{ kJ}$$

FLIP

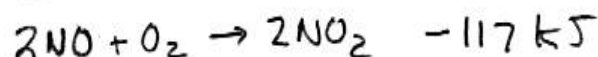


ΔH°



$$\Delta H^\circ_{\text{rxn}} = 33 \text{ kJ}$$

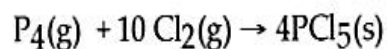
x 2



Flip (reverse) the first equation and multiply the second equation by 2 then add both equations

38.

Use the standard reaction enthalpies given below to determine $\Delta H^\circ_{\text{rxn}}$ for the following reaction:

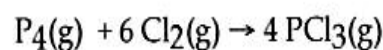


$$\Delta H^\circ_{\text{rxn}} = ?$$

Given:



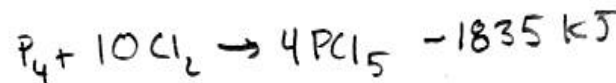
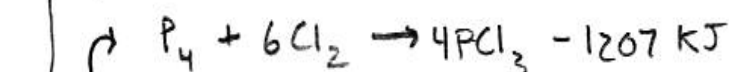
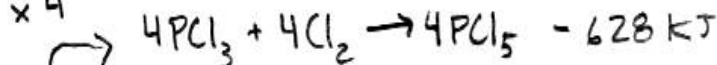
$$\Delta H^\circ_{\text{rxn}} = 157 \text{ kJ}$$



$$\Delta H^\circ_{\text{rxn}} = -1207 \text{ kJ}$$

FLIP

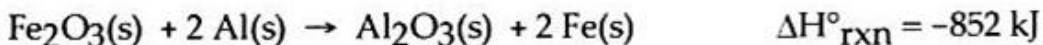
x 4



Flip (reverse) the first equation and multiply it by 4 then added to the second equation

39.

How much energy is released during the formation of 98.7 g of Fe, according to the reaction below?



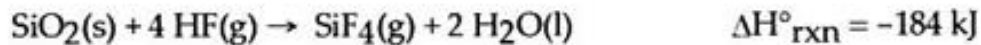
$$98.7 \text{ g} \times \frac{\text{mol Fe}}{55.847 \text{ g}} \times \frac{852 \text{ kJ}}{2 \text{ mol Fe}} = 753 \text{ kJ}$$

Solution:

<u>2 Fe</u>	<u>ΔH</u>
2 x 55.847 g	-852 kJ
98.7 g	? kJ

40.

Using the following information, what mass of HF must react in order to produce 345 kJ of energy? Assume excess SiO_2 .



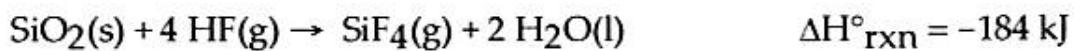
$$345 \text{ kJ} \times \frac{4 \text{ mol HF}}{184 \text{ kJ}} \times \frac{20.006 \text{ g}}{\text{mol HF}} = 1.50 \times 10^2 \text{ g}$$

Solution:

<u>4 HF</u>	<u>ΔH</u>
4 x 20 g	-184 kJ
?	-345 kJ

41.

Using the following information, what mass of H₂O must form in order to produce 975 kJ of energy?



$$975 \text{ kJ} \times \frac{2 \text{ mol H}_2\text{O}}{184 \text{ kJ}} \times \frac{18.016 \text{ g}}{\text{mol H}_2\text{O}} = 191 \text{ g}$$

Solution:

<u>2H₂O</u>	<u>ΔH</u>
2 x 18 g	-184 kJ
?	-975 kJ
