



Exam 4 Question Types & Practice Exam 4

1 definitions and characteristics of systems (easy)

There are three types of systems:

open  heat + matter exchanged (we are open, earth is open)

closed  heat exchanged (coffee cup with lid, bomb calorimeter)

isolated  heat and matter retained (thermos or universe)

know these definitions. period

001 10.0 points

An isolated system allows for the flow of...?

1. none of these correct
2. sound waves
3. kinetic energy
4. matter
5. heat

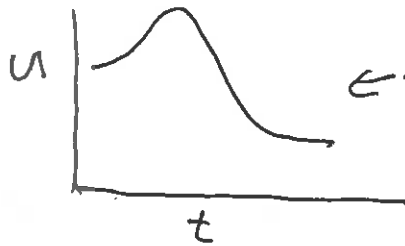
Explanation:

In the natural sciences an isolated system is a physical system without any external exchange - neither matter nor energy (as heat or work) can enter or exit, but can only move around inside.

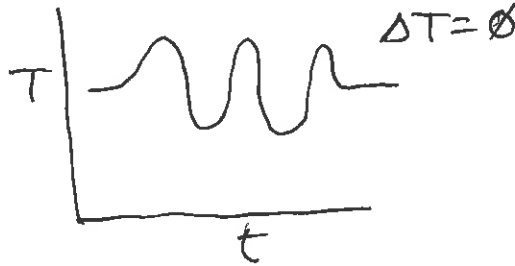
2 definitions and characteristics of state functions medium (tricky)

State functions care about (how it ends - how it starts)

They don't care about path: $\Delta(G, H, T, V, P, n, S, U)$



$\leftarrow \Delta U \text{ is } \ominus$



If you remember the definition and don't freak out the answer is easy

002 10.0 points

Which of the following is true of a general thermodynamic state function?

1. The change in the value of a state function is always negative for a spontaneous reaction.
2. The value of the state function remains constant.
3. The value of a state function does NOT change with a change in temperature of a process.
4. The change of the value of a state function is independent of the path of a process.
correct
5. The change in the value of the state function is always positive for endothermic processes.

Explanation:

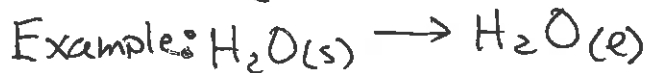
A change in a state function describes a difference between the two states. It is independent of the process or pathway by which the change occurs.

3 thermodynamic conventions - signs/descriptions for q and w medium (Nicky)

"Be the system"



you will be given a chemical or physical process and asked sign of q + w

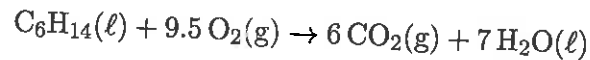


read the question carefully!

$q \equiv -$ because heat enters system
 $w = 0$ because $\Delta V = 0$

003 10.0 points

Consider the reaction



at constant pressure. Which response is true?

1. No work is done as the reaction occurs.
2. Work is done by the system as the reaction occurs.
3. Work is done on the system as the reaction occurs. **correct**
4. Work may be done on or by the system as the reaction occurs, depending upon the temperature.

Explanation:

$n_i = 9.5 \text{ mol gas}$

$n_f = 6 \text{ mol gas}$

For $P = \text{const}$,

$$w = -P \Delta V = -(\Delta n) RT,$$

$$\Delta n = n_f - n_i = -3.5 \text{ mol gas.}$$

For $-\Delta n$, w will be positive, which indicates that work was done on the system, which progresses as the reaction progresses.

4

heat capacity, specific heat capacity theory/application medium

By definition $\Delta H = C \Delta T$ so $C = \frac{\Delta H}{\Delta T}$

- heat capacity is measure of ability to absorb heat into internal motion

- C is big in H_2O with vibrations, H-bonds $C = 4 \text{ J/g}^\circ\text{C}$

- specific heat is $\Delta H = mC\Delta T$ $C = \frac{\Delta H}{m\Delta T}$ where mass (amount) matters

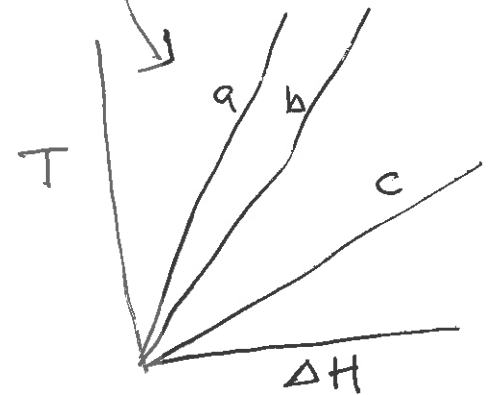
- H_2O has large C , metals are 10 times less

004 10.0 points

Consider the following specific heats: copper, $0.384 \text{ J/g}^\circ\text{C}$; lead, $0.159 \text{ J/g}^\circ\text{C}$; water, $4.18 \text{ J/g}^\circ\text{C}$; glass, $0.502 \text{ J/g}^\circ\text{C}$. If the same amount of heat is added to identical masses of each of these substances, which substance attains the highest temperature? (Assume that they all have the same initial temperature.)

1. copper
2. water
3. lead correct
4. glass

Hint, study plots of T vs ΔH



what is difference between a, b, c heat capacities?

- 5 internal energy, enthalpy definitions & 1st law theory medium
 you will be given several T/F statements about ΔU and ΔH
- 1st Law $\Delta U_{univ} = \Delta U_{sys} + \Delta U_{surv} = 0$ energy of universe conserved
 - $\Delta U_{sys} = -\Delta U_{surv}$ from first law is basis of calorimetry
 - $\Delta H = \Delta U + P\Delta V$ to permit measure of q in open systems
 - $\Delta H = q_p$ $\Delta U = q_v$
 ↑ ↑
 atm closed volume
 - ΔH and ΔU are usually similar magnitudes because $P\Delta V$ is often small or 0 if ΔV is small or 0

005 10.0 points

Which of the following statements is/are true concerning the first law of thermodynamics?

- The internal energy (U) of the universe is conserved.
- The internal energy of a system plus that of its surroundings is conserved.
- The change in internal energy (ΔU) of a system and its surroundings can have the same sign.

1. I, II correct

2. III only

3. II, III

4. I only

5. I, III

6. II only

7. I, II, III

Explanation:

Statement I and II are true; the first law states that the internal energy of the universe is conserved and since the system plus the surroundings is the universe, their sum is also conserved. Statement III is false; for example, if both the system and its surroundings had a positive change in internal energy, then the internal energy of the universe would have increased or decreased - in violation of the first law.

6 work calculation (gases) medium

$$W = -P_{\text{ext}} \Delta V$$

This is a plug and chug. Only difficulty is units. Work is typically kJ but $P\Delta V$ is

l atm so need to use conversion $1 \text{ l atm} = 101.3 \text{ kJ}$

★ ★ signs!! if V decreases, work on system is \oplus
if V increases, work on surroundings is \ominus

be able to do practice exam question

006 10.0 points

1.95 mol of an ideal gas at 300 K and 3.00 atm expands from 16 L to 28 L and a final pressure of 1.20 atm in two steps:

(1) the gas is cooled at constant volume until its pressure has fallen to 1.20 atm, and

(2) it is heated and allowed to expand against a constant pressure of 1.20 atm un-

til its volume reaches 28 L.

Which of the following is CORRECT?

1. $w = 0$ for the overall process
 2. $w = -6.03 \text{ kJ}$ for the overall process
 3. $w = -4.57 \text{ kJ}$ for (1) and $w = -1.46 \text{ kJ}$ for (2)
 4. $w = 0$ for (1) and $w = -1.46 \text{ kJ}$ for (2)
- correct

5. $w = -4.57 \text{ kJ}$ for the overall process

Explanation:

For step (1): If there is no change in volume, $w = 0$.

For step (2): For expansion against a constant external pressure,

$$\begin{aligned} w &= -P_{\text{ext}} \Delta V \\ &= (-1.2 \text{ atm})(18 \text{ L} - 6 \text{ L}) \\ &\quad \times (101.325 \text{ J} \cdot \text{L}^{-1} \cdot \text{atm}^{-1}) \\ &= -1.45908 \text{ kJ.} \end{aligned}$$

The total work done by the system would be the sum of the work for each step.

7

 $\Delta U = q + w$ calculation easySimple plug and chug. Given 2 or 3 ($\Delta U, q, w$) solve for unknown

Only issues
signs. q or w on
system is \oplus , q or w
on surroundings is \ominus

007 10.0 points

A CD player and its battery together do 500 kJ of work, and the battery also releases 250 kJ of energy as heat and the CD player releases 50 kJ as heat due to friction from spinning. What is the change in internal energy of the system, with the system regarded as the battery and CD player together?

1. -800 kJ correct
2. +200 kJ
3. -200 kJ
4. -700 kJ
5. -750 kJ

Explanation:

Heat *from* the CD player is -50 kJ.

Heat *from* the battery is -500 kJ.

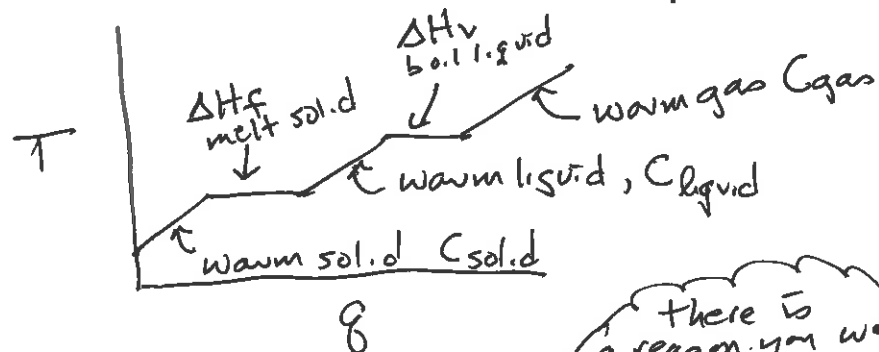
Work *from both together* on the surroundings is -250 kJ.

This question is testing your ability to see what the system is, and then look at **ONLY** the energy flow for the system. Here the system is the battery and the CD player together.

$$\begin{aligned}\Delta U &= q + w \\ &= [-50 \text{ kJ} + (-250 \text{ kJ})] + (-500 \text{ kJ}) \\ &= -800 \text{ kJ}\end{aligned}$$

8

heat flow across phase changes calculation medium



Given starting and ending T_i ,
have to do between 1 and 5
calculations, then add q values up.

Example $50^\circ \rightarrow 150^\circ$ 3 calculations
for H_2O

$0^\circ \rightarrow 50^\circ$ 2 calculations

there is
a reason you were
given H_2O 's values on coversheet!!

calculations use either

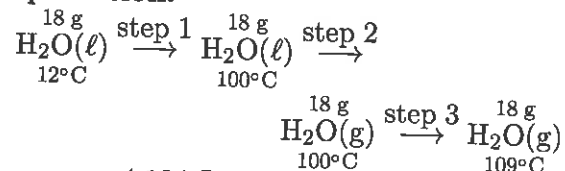
$q = mc\Delta T$ for warming

or $q = (\text{mass})(\Delta H_{\text{v or f}})$
for phase change

008 10.0 points

The specific heat of liquid water is $4.184 \text{ J/g}\cdot^\circ\text{C}$, and of steam $2.03 \text{ J/g}\cdot^\circ\text{C}$. The heat of vaporization of water (ℓ) is 2.26 kJ/g and its boiling point is 100°C . What is the total heat flow when 18 grams of water at 12°C are heated to become steam at 109°C ?

1. 44.4 kJ
2. 47.6 kJ correct
3. under 28 kJ
4. 48.9 kJ
5. over 55 kJ
6. 31.7 kJ
7. 40.7 kJ

Explanation:

$$\text{Step 1: } \frac{4.184 \text{ J}}{\text{g}\cdot^\circ\text{C}} \cdot (18 \text{ g}) \cdot (100 - 12)^\circ\text{C} = 6,627 \text{ J}$$

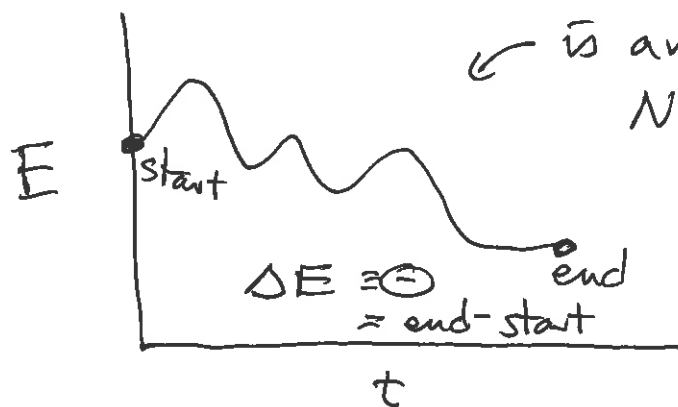
$$\text{Step 2: } \frac{2.26 \text{ kJ}}{\text{g}} \cdot (18 \text{ g}) \cdot \frac{1000 \text{ J}}{1 \text{ kJ}} = 40,680 \text{ J}$$

$$\text{Step 3: } \frac{2.03 \text{ J}}{\text{g}\cdot^\circ\text{C}} \cdot (18 \text{ g}) \cdot (109 - 100)^\circ\text{C} = 329 \text{ J}$$

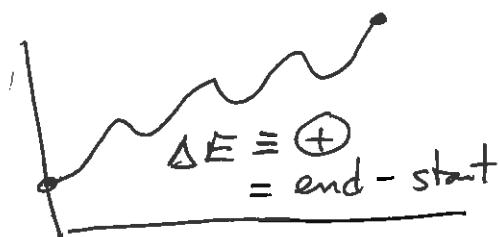
$$\text{Total} = 6627 \text{ J} + 329 \text{ J} + 40,680 \text{ J} = 47,636 \text{ J} = 47.636 \text{ kJ}$$

9

interpreting potential energy (Rxn coordinate) diagrams *easy*

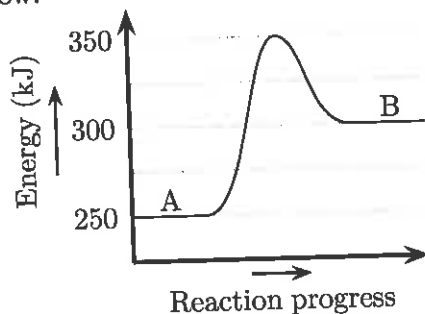


← is an energy profile that is E to surroundings
 Note we don't care about the *wavy* path
 till kinetics in the spring



009 10.0 points

Refer to the potential energy diagram shown below.



What is the change in enthalpy (ΔH) for the reaction



1. +50 kJ, endothermic **correct**
2. -50 kJ, exothermic
3. +50 kJ, exothermic
4. -250 kJ, endothermic
5. -150 kJ, endothermic

Explanation:

$$\Delta H_i = \Delta H_A = 250 \text{ kJ}$$

$$\Delta H_f = \Delta H_B = 300 \text{ kJ}$$

$$\begin{aligned} \Delta H &= \Delta H_f - \Delta H_i \\ &= 300 \text{ kJ} - 250 \text{ kJ} \\ &= 50 \text{ kJ} \end{aligned}$$

Notice that ΔH is positive and therefore the reaction is endothermic.

10 calorimetry theory medium

Be able to explain how a calorimeter works and the definition of each variable in q calculation



$$\Delta U_{\text{sys}} = \Delta U_{\text{cheetos}} = mC\Delta T_{\text{cheetos}} = -\Delta U_{\text{sur}} = - \left(mC\Delta T_{\text{water}} + C\Delta T_{\text{calorimeter}} \right)$$

- more water, smaller ΔT
- large $C_{\text{H}_2\text{O}}$, smaller ΔT

• $C_{\text{calorimeter}}$ is heat capacity in J/°C $C_{\text{H}_2\text{O}}$ is specific heat in J/g°C

Which of the following is NOT a feature of the bomb calorimetry apparatus used to measure the internal energy of a reaction?

1. The heat capacity of the calorimeter should be known to accurately correct for any heat lost to it.
2. The thermometer is inserted directly into the reaction vessel to measure ΔT of the reaction. **correct**
3. Large quantities of water surrounding the reaction vessel absorb the majority of the heat loss.
4. The volume of the reaction vessel is held constant to eliminate energy released as work.
5. The large heat capacity of water is beneficial in measuring heat released by combustion reactions.

Explanation:

The thermometer is placed in the water that surrounds the reaction vessel

11 calorimetry calculation medium

$$\Delta U_{\text{sys}} = +\Delta U_{\text{surv}} = - \left(m_{\text{H}_2\text{O}} C_{\text{H}_2\text{O}} \Delta T + C_{\text{calorimeter}} \Delta T \right)$$

- easy plug and chug using ΔT twice
- two $C_{\text{H}_2\text{O}}$, $C_{\text{calorimeter}}$ values
- convert l \rightarrow ml to grams
- usual answer is per amount like J/cheetos or kJ/gram

Correct for per like explodes 4 cheetos but 8 per cheetos

011 10.0 points

A 1.00 g sample of *n*-hexane (C_6H_{14}) undergoes complete combustion with excess O_2 in a bomb calorimeter. The temperature of the 1502 g of water surrounding the bomb rises from 22.64°C to 29.30°C . The heat capacity of the hardware component of the calorimeter (everything that is not water) is $4042 \text{ J}/^\circ\text{C}$. What is ΔU for the combustion of *n*- C_6H_{14} ? One mole of *n*- C_6H_{14} is 86.1 g. The specific heat of water is $4.184 \text{ J}/\text{g}\cdot^\circ\text{C}$.

1. $-1.15 \times 10^4 \text{ kJ/mol}$
2. $-4.52 \times 10^3 \text{ kJ/mol}$
3. $-7.40 \times 10^4 \text{ kJ/mol}$
4. $-5.92 \times 10^3 \text{ kJ/mol}$ correct
5. $-9.96 \times 10^3 \text{ kJ/mol}$

Explanation:

$$\begin{aligned} m_{\text{C}_6\text{H}_8} &= 1.00 \text{ g} & m_{\text{water}} &= 1502 \text{ g} \\ \text{SH} &= 4.184 \text{ J}/\text{g}\cdot^\circ\text{C} & \text{HC} &= 4042 \text{ J}/^\circ\text{C} \\ \Delta T &= 29.30^\circ\text{C} - 22.64^\circ\text{C} = 6.66^\circ\text{C} \end{aligned}$$

The increase in the water temperature is $29.30^\circ\text{C} - 22.64^\circ\text{C} = 6.66^\circ\text{C}$. The amount of heat responsible for this increase in temperature for 1502 g of water is

$$\begin{aligned} q &= (6.66^\circ\text{C}) \left(4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (1502 \text{ g}) \\ &= 41854 \text{ J} = 41.85 \text{ kJ} \end{aligned}$$

The amount of heat responsible for the warming of the calorimeter is

$$\begin{aligned} q &= (6.66^\circ\text{C})(4042 \text{ J}/^\circ\text{C}) \\ &= 26920 \text{ J} = 26.92 \text{ kJ} \end{aligned}$$

The amount of heat released on the reaction is thus $41.85 \text{ kJ} + 26.92 \text{ kJ} = 68.77 \text{ kJ}$ per g of *n*-hexane.

Per mol of *n*-hexane, this becomes

$$\left(68.77 \frac{\text{kJ}}{\text{g}} \right) \left(86.1 \frac{\text{g}}{\text{mol}} \right) = 5921 \text{ kJ/mol}$$

12 formation constants definition easy

"Elements in their Standard States"

↓
H, C, O, N

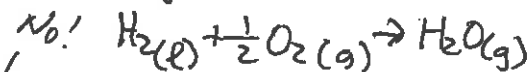
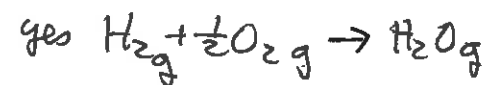
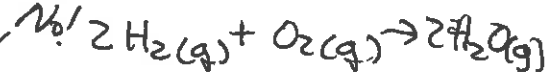
↓
room temp
and pressure

↓
H₂(gas)
C (graphite)
Hg (liquid)

are combined to form a compound (1 mole)

which of a series of reactions is a formation reaction

Example



↳ H₂(l) is a NoNo
2 moles product is NoNo

012 10.0 points

Which of the reactions below is a formation reaction?

1. $\text{H}_2(\text{g}) + \text{C}_{\text{graphite}}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
2. $2\text{H}_2(\text{g}) + 2\text{C}_{\text{graphite}}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{CH}_2\text{O}(\text{l})$
3. $2\text{Fe}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s})$ correct
4. $\text{N}_2(\text{l}) + 2\text{H}_2(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{l})$

Explanation:

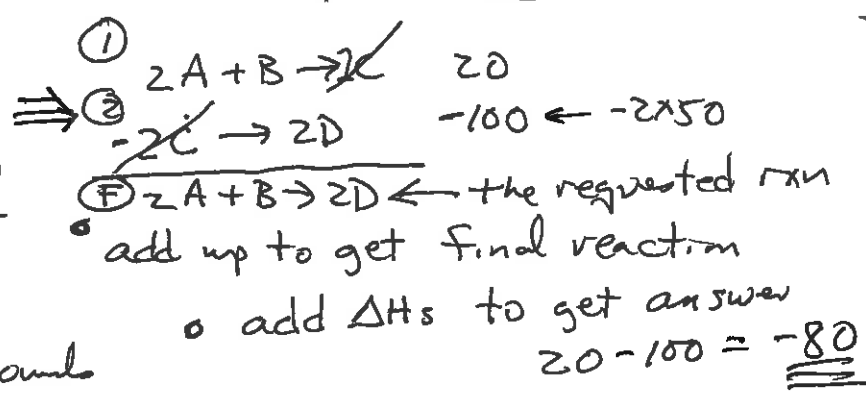
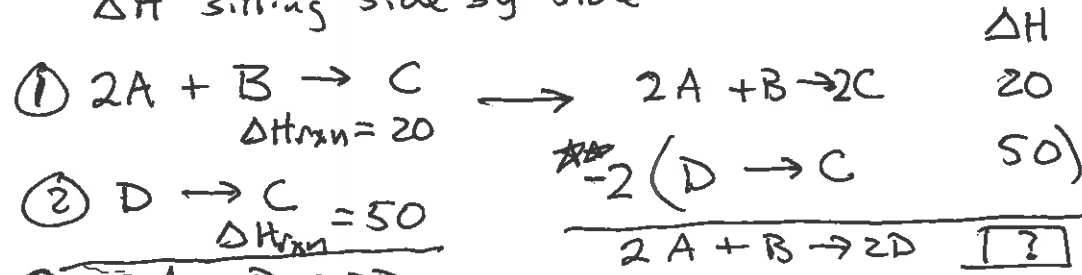
A formation reaction produces exactly one mole of one product from elements in their standard states.

13

Hess' law - Rxn enthalpies hard

this is hard because it is messy. Rewrite the problem to clean it up with reaction and ΔH sitting side by side

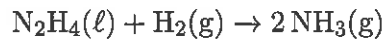
Hess' Law says you can combine various reactions and cancel them out to add up $\Delta H_{rxn1} + \Delta H_{rxn2} = \Delta H_{rxn\ total}$



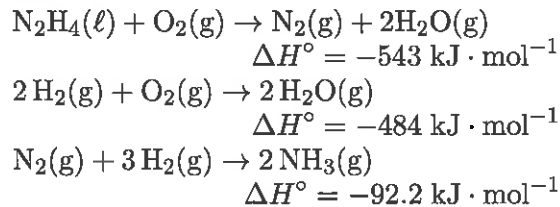
Then multiply through with +/- and coefficients to cancel all but final compounds

The desired rxn Find

Calculate the standard reaction enthalpy for the reaction



given

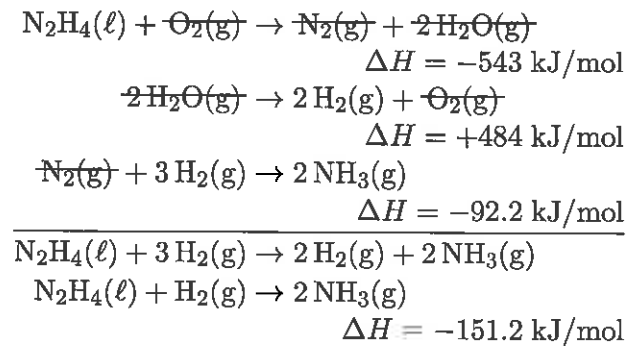


Messy \rightarrow rewrite to organize

1. $-1119 \text{ kJ} \cdot \text{mol}^{-1}$
2. $-151 \text{ kJ} \cdot \text{mol}^{-1}$ correct
3. $-935 \text{ kJ} \cdot \text{mol}^{-1}$
4. $-243 \text{ kJ} \cdot \text{mol}^{-1}$
5. $-59 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:

We need to reverse the second reaction and add them:



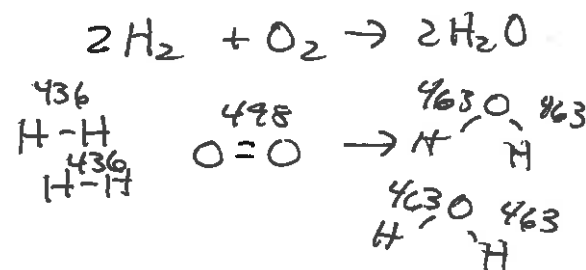
this one is easy. Just one change, Exam is more than one switch. But usually +/- a small integer

14 Hess' law - bond energy *easy but messy*

Give table of BE on front page

- write the Lewis structures bond by bond
- label each bond with B.E.

- add up reactants $436 \times 2 + 498$ and products 4×463
- $\sum \text{reactants} - \sum \text{products} = \text{BE} \hat{=} -500$ (actual is -482)
- Don't over round off, Take your time with math

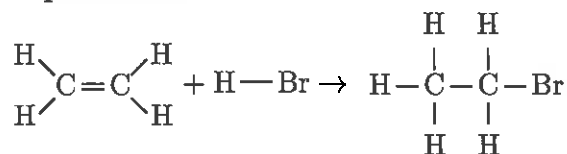


014 10.0 points

Estimate the heat released when ethene ($\text{CH}_2=\text{CH}_2$) reacts with HBr to give $\text{CH}_3\text{CH}_2\text{Br}$.

1. 200 kJ/mol
2. 76 kJ/mol correct
3. 1036 kJ/mol
4. 424 kJ/mol
5. 470 kJ/mol

Explanation:



$$\begin{aligned} \Delta H &= \sum E_{\text{break}} - \sum E_{\text{make}} \\ &= [(\text{C}=\text{C}) + (\text{H}-\text{Br})] \\ &\quad - [(\text{C}-\text{H}) + (\text{C}-\text{Br}) + (\text{C}-\text{C})] \\ &= (602 \text{ kJ/mol} + 366 \text{ kJ/mol}) \\ &\quad - [413 \text{ kJ/mol} + 285 \text{ kJ/mol} \\ &\quad \quad + 346 \text{ kJ/mol}] \\ &= -76 \text{ kJ/mol}, \end{aligned}$$

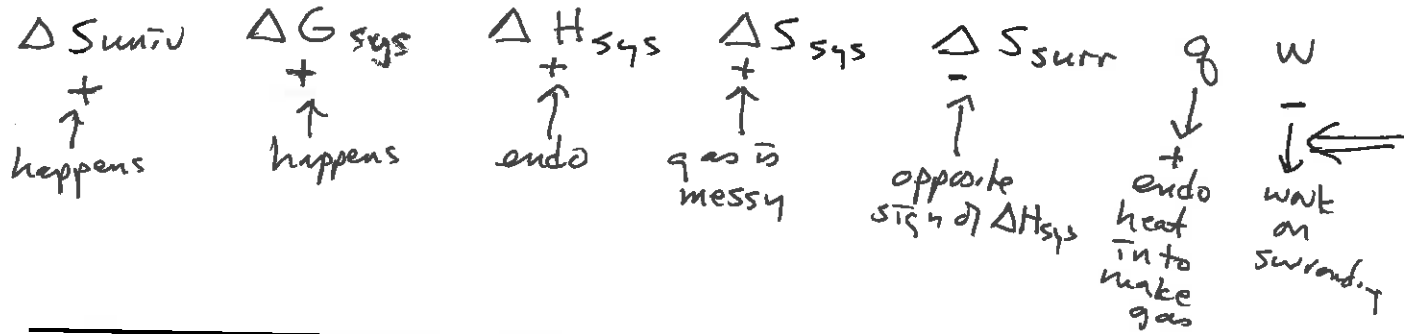
so 76 kJ/mol of heat was released.

15

thermodynamic conventions - signs/descriptions for ΔG , ΔH and ΔS ^{medium}

For a given set of conditions be able to assign

be the system



For example if given a phase change producing gas at high T, then the signs are as shown starting with $\Delta H_{sys} \oplus, \Delta S_{sys} \oplus$ because that is clue given. And $w \ominus$ because gas formed in clue

015 10.0 points

Consider a reaction that is non-spontaneous at all temperatures. What would be the signs of ΔG_{sys} , ΔH_{surr} , and ΔS_{univ} respectively for such a reaction?

1. +, -, +
2. -, +, -
3. +, -, - correct
4. -, +, +
5. +, +, +

Explanation:

For a reaction that is non-spontaneous at all temperatures, the free energy of the system will increase and the entropy of the universe will decrease. Such a reaction must be endothermic, and the heat it gains will be lost from the surroundings.

16 ranking molar entropy *easy*

Given a collection of compounds, rank S_m° using logic:

$s \rightarrow l \rightarrow g$ is larger S_m°

small \rightarrow large is larger S_m°

simple \rightarrow complicated is larger S_m°

Examples
are in the
cover page
tables

this is
not ΔS
this is
for cmpds

Example $H_2O(l) < H_2O_2(l) < H_2O(g) < H_2O_2(g)$

Which one shows the substances in the decreasing order of their molar entropy?

1. C(s), $H_2O(l)$, $H_2O(g)$, $H_2O(s)$
2. $H_2O(g)$, $H_2O(l)$, $H_2O(s)$, C(s) **correct**
3. None of these
4. $H_2O(s)$, $H_2O(l)$, $H_2O(g)$, C(s)
5. C(s), $H_2O(g)$, $H_2O(l)$, $H_2O(s)$
6. C(s), $H_2O(s)$, $H_2O(l)$, $H_2O(g)$

Explanation:

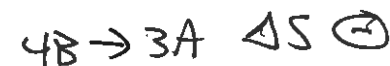
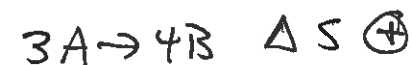
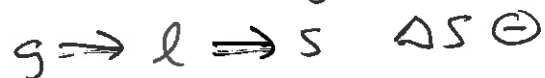
Gases will have a higher entropy than liquids so we expect $H_2O(l)$ to have the lowest molar entropy. The gases will increase in entropy in the order $Ne(g) < Ar(g) < CO_2(g)$. Ne and Ar are both atoms so they should have less entropy than a molecular substance, which has more complexity. Ar will have a higher entropy than Ne because it has a larger mass and more fundamental particles.

The correct order is

$H_2O(l) < Ne(g) < Ar(g) < CO_2(g)$.

17 predicting changes in entropy easy

Given a series of chemical and physical processes, predict sign of ΔS_{sys}



017 10.0 points

Consider the following processes. (Treat all gases as ideal.)

- I) The pressure of one mole of oxygen gas is allowed to double isothermally.
- II) Carbon dioxide is allowed to expand isothermally to 10 times its original volume.
- III) The temperature of one mole of helium is increased 25°C at constant pressure.
- IV) Nitrogen gas is compressed isothermally to one half its original volume.
- V) A glass of water loses 100 J of energy reversibly at 30°C .

Which of these processes leads to an increase in entropy?

1. III and V
2. I and II
3. I and IV
4. V
5. II and III correct

18 stat thermo positional entropy medium

$$S = k \ln W = k \ln \#^N \quad \# \Rightarrow \text{orientations} \quad N \Rightarrow \text{\# of molecules}$$

= 2 for CO

= 3 for BH₂Cl

= 4 for CH₃Cl

= 6 for C₆H₅Cl

• remember $k N_{\text{moles}} = R = 8.3$

• remember $\ln 2 = 0.7 \quad \ln 3 = 1.1$

• remember $\ln x^y = y \ln x \quad \ln xy = \ln x + \ln y$

Example

What is S for 5 moles of CH₃Cl

$$S = k \ln 4^{5N} = k \ln 2^{2 \times 5N}$$

$$\downarrow$$

$$= 10 N k \ln 2 = 10 (8.3) \ln 2$$

$$\rightarrow 83 (0.7) \approx 60 \frac{\text{J}}{\text{K}}$$

What is the entropy at $T = 0 \text{ K}$ for one mole of chloroform (CHCl₃)?

1. $1.9 \times 10^{-23} \text{ J K}^{-1}$
2. 1.38 J K^{-1}
3. 11.5 J K^{-1} correct
4. -11.5 J K^{-1}
5. 0 J K^{-1}

Explanation:

Since the question is concerned with the residual entropy of a mole of chloroform, which has 4 orientations, we can use the Boltzmann equation to calculate the residual entropy.

$$S = N_a k \ln W = R \ln 4 = 11.5 \text{ J K}^{-1}$$

20

second law theory medium

Here are some maths

$$\bullet \Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{sur}} > 0 \text{ happens}$$

$$\bullet \Delta S = \frac{q}{T} \Leftarrow \text{which means } \Delta S \text{ is larger for high } q + \text{ low } T$$

so exploding H_2 balloon at north pole is big ΔS

$$\bullet \Delta G_{\text{sys}} = -T \Delta S_{\text{univ}} \text{ is how we get to use } \Delta G_{\text{sys}} \text{ rather than } \Delta S_{\text{univ}}$$

$$\bullet \Delta S_{\text{sur}} = -\frac{\Delta H_{\text{sys}}}{T} \text{ is how heat from exothermic reaction makes chaos in surroundings}$$

020 10.0 points

Advertising claims sometimes state that adding something mechanical to a car's engine will allow it to recover 100% of the energy that comes from burning gasoline. You should be skeptical of such claims because they violate the

1. first law of thermodynamics.
2. activation energy requirements of all chemical reactions.
3. law of conservation of matter.
4. second law of thermodynamics. **correct**

Explanation:

If you burn gasoline in an engine to move a car, you are ultimately converting chemical potential energy into kinetic energy. But much of this energy is lost as heat. There is **NO** way to make any energy conversion 100% efficient.

21

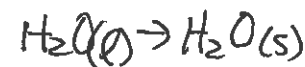
second law calculation (ΔS of surroundings) mediumIn calculations ΔS_{univ} , ΔS_{sys} comes from table of molar entropies.But ΔS_{surr} comes from table of ΔH_{rxn}

$$\Delta S_{\text{surr}} = -\frac{\Delta H_{\text{sys}}}{T}$$

- make sure signs are right, ΔS_{surr} opposite ΔH_{sys}
- get units right. Use J for everything, not kJ

Example What is ΔS_{surr} if $\Delta S_{\text{sys}} = 22 \text{ J/K}$ and $\Delta H_{\text{rxn}} = 6 \text{ kJ}$ at 200K

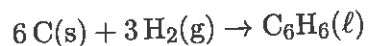
$$\Delta S_{\text{surr}} = -\left(\frac{-6000 \text{ J}}{200}\right) = +30 \frac{\text{J}}{\text{K}}$$



- Note you don't use ΔS_{sys}
- Note $l \rightarrow s$ is exo so use -6000 not +6000
- Note ΔS_{surr} \oplus became exothermic

↑
2 negatives, one from equation and one from sign of water ΔH freezing

021 10.0 points

Calculate $\Delta S_{\text{surr}}^{\circ}$ at 298 K for the reaction

$$\Delta H_{\text{r}}^{\circ} = +49.0 \text{ kJ}\cdot\text{mol}^{-1} \text{ and } \Delta S_{\text{r}}^{\circ} = -253 \text{ J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}.$$

1. $-417 \text{ J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$

2. $-164 \text{ J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$ correct

3. $+253 \text{ J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$

Explanation:

$$\Delta H_{\text{r}}^{\circ} = 49000 \text{ J}\cdot\text{mol}^{-1}$$

$$T = 298 \text{ K}$$

$$\begin{aligned} \Delta S_{\text{surr}}^{\circ} &= \frac{q_{\text{surr}}}{T} = \frac{-q}{T} = \frac{-\Delta H_{\text{r}}^{\circ}}{298 \text{ K}} \\ &= \frac{-(+49000 \text{ J}\cdot\text{mol}^{-1})}{298 \text{ K}} \\ &= -164.43 \text{ J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}. \end{aligned}$$

22 relating reaction spontaneity to T for phase changes *medium*

$$\Delta G = \Delta H - T\Delta S = 0 \text{ at phase change so}$$

$$\Delta H = T\Delta S$$

$$T_{\text{phase}} = \frac{\Delta H}{\Delta S}$$

Example What is T for H₂O boiling if

$$\Delta H_{\text{vap}} = 40 \text{ kJ and } \Delta S_{\text{vap}} = 107 \text{ J/K}$$

$$T = \frac{40000}{107} \approx 373 \Rightarrow 100^\circ\text{C}$$

• be careful with units of T
K vs °C

• Not. cc how similar question 21 is.

022 10.0 points

For the evaporation of water from an open pan at 25°C, the values of ΔS for the water, the surroundings, and the universe must be, respectively,

1. positive, positive, positive.
2. positive, negative, zero.
3. None of these is correct.
4. negative, negative, negative.
5. positive, negative, positive. **correct**

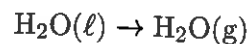
Explanation:

The process is spontaneous, which means $\Delta S_{\text{universe}} > 0$ according to the Second Law of Thermodynamics.

Entropy (S) is high for systems with high degrees of freedom, disorder or randomness and low for systems with low degrees of freedom, disorder or randomness.

$$S(\text{g}) > S(\text{l}) > S(\text{s}).$$

The system is



so $\Delta S_{\text{system}} > 0$.

The entropy of the surroundings must be negative. Energy is removed from the surroundings to get the water to evaporate.

relating reaction spontaneity to T for chemical reactions

$$\Delta G_{sys} = \Delta H_{sys} - T\Delta S_{sys}$$

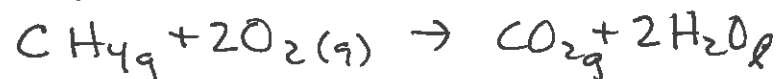
+	+	happens high T
+	-	never happens
-	+	always happens
-	-	happens low T

$$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{sur}$$

+	+	always
+	-	happens high T
-	+	happens low T
-	-	never

For a process, be able to assign
 ΔS_{univ} , ΔS_{sys} , ΔS_{sur} , ΔG_{sys} , ΔH_{sys}

Examples



$\Delta H_{sys} \ominus$ combustion, $\Delta S_{sys} \ominus$ $g \rightarrow l$

so $\Delta S_{sur} \oplus$, $\Delta S_{univ} \oplus$ and $\Delta G_{sys} \ominus$ at low T

Example $H_2O(g) \xrightarrow{90^\circ} H_2O(l)$ happens -

$\Delta G \ominus$, $\Delta S_{univ} \oplus$, $\Delta S_{sys} \ominus$, $\Delta H_{sys} \ominus$, $\Delta S_{sur} \oplus$
 note happens low T ($< 100^\circ C$)

For the reaction



$\Delta H_r^\circ = +198 \text{ kJ}\cdot\text{mol}^{-1}$ at 298 K. Which statement is true for this reaction?

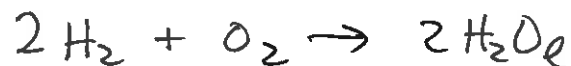
1. The reaction is driven by the enthalpy.
2. The reaction will not be spontaneous at any temperature.
3. The reaction will not be spontaneous at high temperatures.
4. ΔG_r° will be positive at high temperatures.
5. ΔG_r° will be negative at high temperatures. **correct**

Explanation:

$\Delta G = \Delta H - T\Delta S$ is used to predict spontaneity. (ΔG is negative for a spontaneous reaction.) ΔH is positive and T is always positive. For the reaction 2 mol gas \rightarrow 3 mol gas. The more moles of gas, the higher the disorder, so ΔS is positive and $\Delta G = (+) - T(+)$. For ΔG to be negative, T must be large.

24 ΔG_{rxn} from table values calculation *med. un but very messy*

Given a chemical reaction



What is ΔG_{rxn} at 300K

• use $\Delta G = \Delta H - T\Delta S$

↑ from table ↑ from table

• use cover sheet data

• Hint. : Combustion reaction $\Delta G \ominus$

• Hint. $T\Delta S$ is usually small

so $\Delta H \sim \Delta G$

Just saying...

$$\Delta H = \sum \Delta H_{f, \text{prod}}^\circ - \sum \Delta H_{f, \text{rec}}^\circ$$

$$-2(286) - (2(0) + 0)$$

became $\Delta H_f = 0$

$$= -572 \text{ kJ}$$

$$\Delta S = \sum S_{m, \text{prod}}^\circ - \sum S_{m, \text{rec}}^\circ$$

$$= 2(70) - [2(130) + 205]$$

$$140 - 465 = -325$$

$$T\Delta S = 300(-325) = -975$$

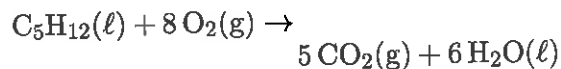
$$\Delta G_{rxn} = -572 - (-975) = -403$$

Awful, just awful.
Can't I use a calculator? NO!!!
But!!

024 10.0 points

What is ΔG_r° for the combustion of liquid *n*-pentane?

1. 3389 kJ/mol
2. -383 kJ/mol
3. 383 kJ/mol
4. -3389 kJ/mol correct
5. -451 kJ/mol
6. 451 kJ/mol



$$\Delta H = [5(-394) + 6(-286)] \text{ kJ/mol}$$

$$- [1(-174) + 8(0)] \text{ kJ/mol}$$

$$= -3512 \text{ kJ/mol}$$

and

$$\Delta S = [5(214) + 6(70)] \text{ J/mol} \cdot \text{K}$$

$$- [1(263) + 8(205)] \text{ J/mol} \cdot \text{K}$$

$$= -413 \text{ J/mol} \cdot \text{K}$$

$$= -0.413 \text{ kJ/mol} \cdot \text{K}$$

$$\Delta G = \Delta H - T\Delta S$$

$$= -3512 \text{ kJ/mol}$$

$$- (298 \text{ K})(-0.413 \text{ kJ/mol} \cdot \text{K})$$

$$= -3389 \text{ kJ/mol}$$

25 predicting compound stability from ΔG_f° easy

If $\Delta G_f^\circ \ominus$ then compound is stable
 If $\Delta G_f^\circ \oplus$ then compound not stable

Diamond as $\Delta G_f^\circ = +3 \text{ kJ/mole}$
 so not stable
 Benzene $\Delta G_f^\circ = +49 \text{ kJ}$
 so not stable
 Ethanol $\Delta G_f^\circ = -278 \text{ kJ}$
 so stable

025 10.0 points

Consider the data below regarding different allotropes of carbon.

carbon allotrope	ΔG_f° kJ · mol ⁻¹
C(s, graphite)	0
C(s, diamond)	2.9
C ₆₀ (s, buckminsterfullerene)	24

Which of the following statements is supported by these data?

1. C₆₀ is thermodynamically more stable than graphite under standard conditions.
2. Graphite could spontaneously form C₆₀ under standard conditions.
3. Formation of graphite from C₆₀ would be exergonic under standard conditions. **correct**
4. Diamond is the least thermodynamically stable allotrope of carbon under standard conditions.