

CHAPTER 10  
ENERGY &  
REACTION SPONTANEITY

10

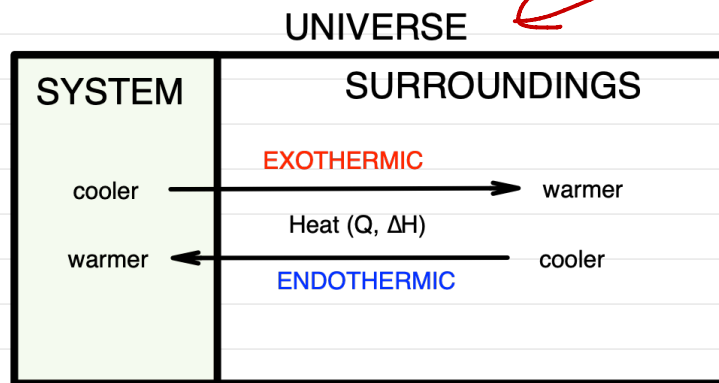
# Chapter 10: Energy

## Thermodynamics

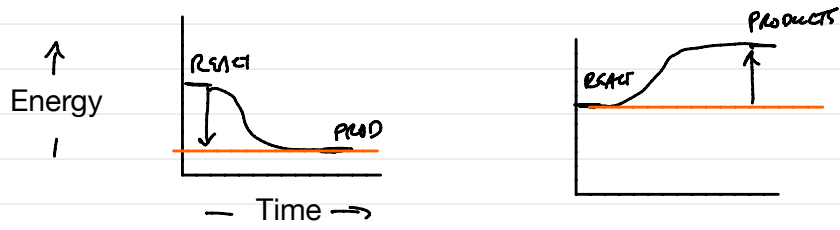
Thermodynamics  
↙ energy/heat  
↘ movement

### Endo- and Exothermic Phenomena

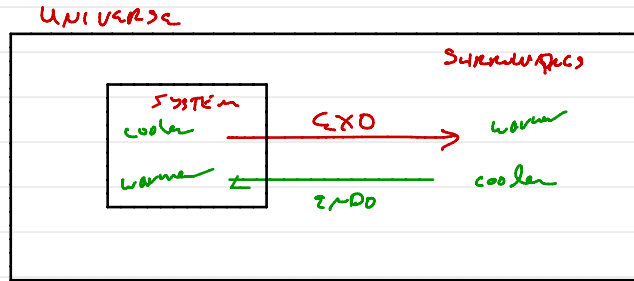
1st Law of Thermodynamics



# Energy Diagrams

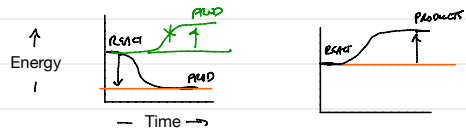


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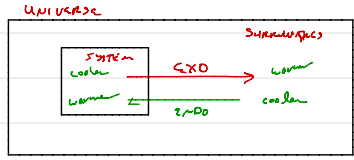


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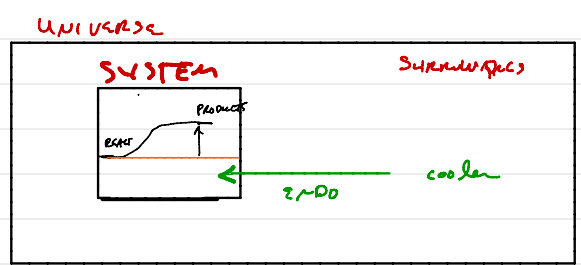
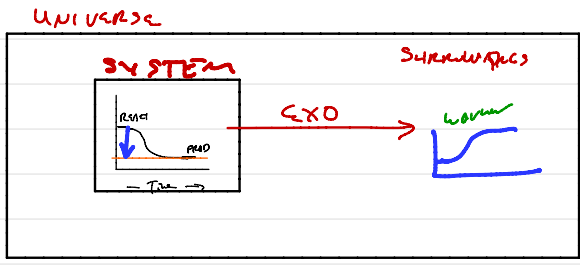
# Energy Diagrams



+



↓



## 2 Types of Energy

- PE - due to POSITION  $\rightarrow$  bond (spring)
- KE - due to MOTION  $\rightarrow$  VRT

## 2 Forms of Energy

(Formal Definition of Energy Change)

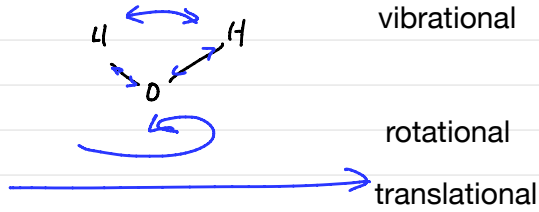
$$\Delta E = Q + W$$

heat  
(camp fire)  $\rightarrow$

$\leftarrow$  work  
(car movement)

¿How do molecules manifest energy?

- ↳ vibrational
- ↳ rotational
- ↳ translational



vibrational

rotational

translational

VRT =  $\left(\frac{1 \text{ kcal}}{1 \text{ Cal}}\right)$   
 $\left(\frac{1000 \text{ cal}}{1 \text{ Cal}}\right)$

$\left(\frac{4.184 \text{ J}}{1 \text{ cal}}\right)$

$\frac{1 \text{ Cal}}{1} = \frac{90 \text{ Cal} \cdot \frac{1000 \text{ cal}}{1 \text{ Cal}}}{1 \text{ Cal}}$

¿What is the difference between HEAT and TEMPERATURE?

heat energy

temperature

$Q = m \cdot s \cdot \Delta T$



$\frac{J}{1} = \frac{J}{g \cdot ^\circ C}$

$\frac{J}{g \cdot ^\circ C}$

(EX) A 1.6 g sample of metal requires 5.8 J of energy to change its temperature from 23 °C to 41 °C. Is the material gold?

$s(\text{Au}) = 0.13 \text{ J/g}\cdot\text{C}$ ;  $s(\text{H}_2\text{O}) = 4.184 \text{ J/g}\cdot\text{C}$

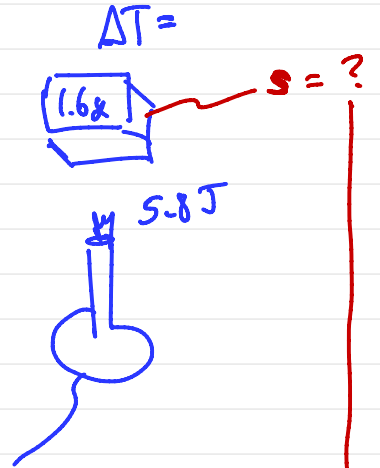
$T_f$

$\Delta T = T_f - T_i$   
 SYS, SUMM  
 $41 - 23 = +18^\circ\text{C}$

heating

~~$\Delta T = T_f - T_i$   
 $23 - 41 = -18^\circ\text{C}$~~

~~cooling~~



SYS	SUPR
$\Delta T$	$\Delta T$



$Q = m \cdot s \cdot \Delta T$   
 $s = \frac{Q}{m \cdot \Delta T} = \frac{5.8 \text{ J}}{1.6 \text{ g} \cdot 18^\circ\text{C}} = 0.20 \frac{\text{J}}{\text{g}\cdot\text{C}}$

$s(\text{Au}) = 0.13$

NOT Gold



(EX) 5.63 g of gold is heated from  $21^{\circ}\text{C}$  to  $32^{\circ}\text{C}$ . How much energy (in calories) is required?

$0.13 \text{ J/g}^{\circ}\text{C}$

$T_i$   $T_f$   $\Delta T = 32 - 21 = 11^{\circ}\text{C}$

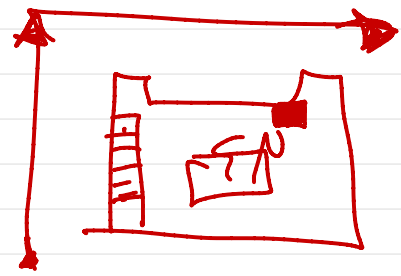
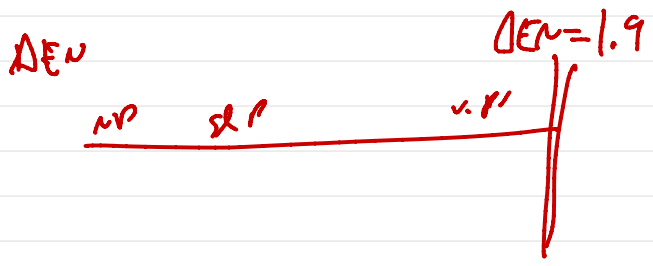
$$\frac{\square \text{ cal}}{1} = \frac{5.63 \text{ g}}{1} \cdot \frac{0.13 \text{ J}}{\text{g}^{\circ}\text{C}} \cdot \frac{11^{\circ}\text{C}}{1} \cdot \frac{1 \text{ cal}}{4.184 \text{ J}}$$

8.1 J, or 1.9 cal

$$Q = m \cdot c \cdot \Delta T$$

$$= (5.63)(0.13)(11)$$

$$=$$





## Free Energy ( $\Delta G$ )

$$\Delta E = Q + W$$
$$\Delta G = \Delta H - T\Delta S = \Delta H + (-T\Delta S)$$

$Q = m \cdot c \cdot \Delta T$

Gibbs Free Energy ("total" energy) = Enthalpy ("Heat" energy) "Q" - Temperature Entropy

TOTAL ENERGY = HEAT ENERGY + WORK ENERGY

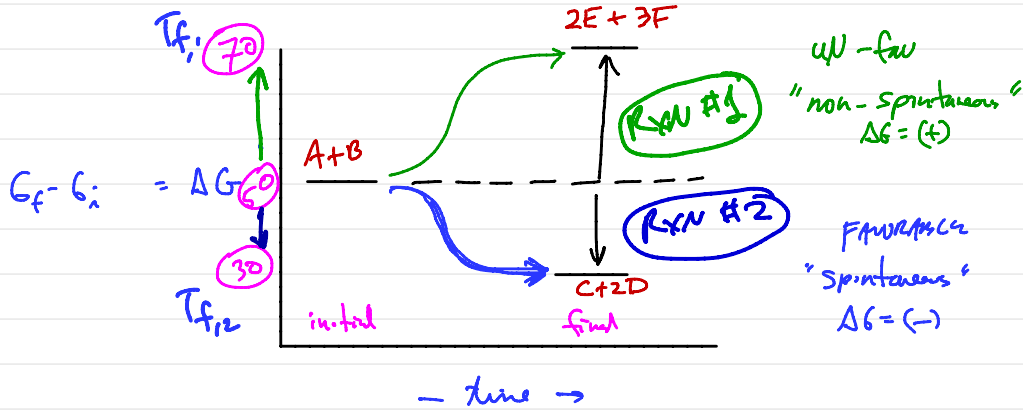
+	+	=	+
-	-	=	+
-	+	=	-
+	-	=	-

## Three Thermodynamic Parameters

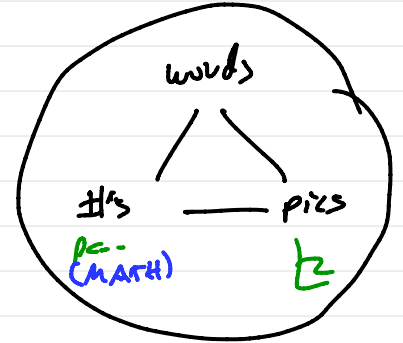
- ↳ Free energy ("total energy")
- ↳ Enthalpy ("heat energy")
- ↳ Entropy ("ordering energy")

Mon, Nov 18

# Energy Diagrams ("picture speak")



KEY TIP:



If  $\Delta G$  is negative, then reaction is "Spontaneous"

$$\Delta G = \Delta G_f - \Delta G_i$$

$$\Delta G_{\text{Rxn 1}} = 70 - 50 = +20$$

$$\Delta G_{\text{Rxn 2}} = 30 - 50 = -20$$

# ΔG Scenarios

# Summary

Key question: Is ΔG negative?

Favorable if...

↳ ΔG = (-)

↳ ΔH = (-) *exo*

↳ ΔS = (+)

Unfavorable if...

↳ ΔG = (+)

↳ ΔH = (+) *endo*

↳ ΔS = (-)

↳ If so, then reaction is Spontaneous. ΔG = (-)

↳ If not, the reaction is not Spontaneous. ΔG = (+)

- (00 = 150 - 250)

$$\Delta G = \Delta H - T\Delta S = \Delta H + (-T\Delta S)$$

#1 (-) = (-) - [T (+)]  
*favourable* = *favourable* - *favourable*

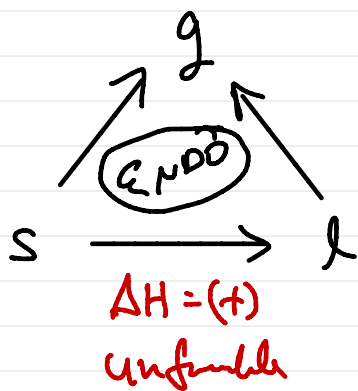
#2 (+) = (+) - [T (-)]  
*unfavourable* = *unfavourable* - *unfavourable*

#3 (?) = (-) - [T (-)]  
*"depeck"* = *favourable* - *unfavourable*

ΔH = (-)  
 Exotherm

#4 (?) = (+) - [T (+)]  
*"depeck"* = *unfavourable* - *favourable*  
 HENT  
 ENTROPY

## Enthalpy of Exo- and Endothermic Reactions



Recall, UNfavorable if...

↳  $\Delta G = (+)$

↳  $\Delta H = (+)$  aka ENDOthermic

↳  $\Delta S = (-)$

## Entropy (measure of "disorder")

- ↳ Entropy is given by symbol  $\Delta S$
- ↳ when  $\Delta S$  is multiplied by T, it become an energy term (" $T\Delta S$ ")
- ↳ the  $T\Delta S$  energy term shows up in equation:  $\Delta G = \Delta H - T\Delta S$
- ↳ it is taken as a measure of Disorder

ENTROPY IS A MEASURE OF . . .

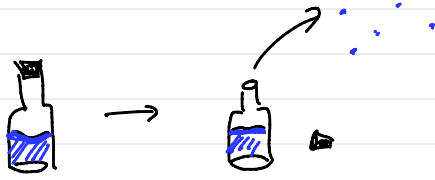
- ↳ disorder
- ↳ randomness
- ↳ freedom
- ↳ chaos
- ↳ run-down
- ↳ energy dispersal
- ↳ conversion of usable energy to UNusable energy

$+ \Delta S$   
more  
d-disorder

The universe tends toward Disorder; toward increasing Entropy

If  $\Delta S(\text{universe})$  is positive, then reaction is "Spontaneous"

## Entropy (measure of “disorder”)



### The Perfume Bottle Problem

#### DILEMMA

For you to be able to smell perfume, the molecules must first vaporize from the liquid to the gas phase, then travel across space and arrive at your nose.

So, what's the problem?

- vaporization ( $l \rightarrow g$ ) is an ENDO process
- vaporization is an UNfavorable change
- so, vaporization does not occur spontaneously (naturally), so, the molecules shouldn't go into the gas phase, and so shouldn't travel to your nose, and so you shouldn't smell it ... BUT YOU DO!

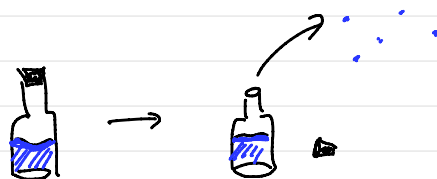
#### REMEDY

- $\Delta H = (+)$  may be UNfavorable (endothermic) for the perfume process, but  $\Delta S = (+)$  is Favorable
- Upon vaporization, the perfume molecules experience a tremendous gain in freedom (or disorder) — they are now free to move “about the cabin,” rather than restricted to the confines of a small, capped bottle.
- The favorable gain in Entropy must have more than enough to offset the UNfavorable endothermic process, such that the overall energy change was favorable.

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

net favorable      UNfavorable (endothermic)      very Favorable (large gain in disorder)

## Entropy (measure of “disorder”)



### The Perfume Bottle Problem

#### DILEMMA

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...→

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*Wed, Nov 20<sup>th</sup>*



## First and Second Laws of Thermodynamics

- ↳ immutable
- ↳ no exceptions

“If your theory is found to be against the second Law of thermodynamics, I give you no hope; there is nothing for it but to collapse in deepest humiliation.

— Arthur Eddington

$$\Sigma = mc^2$$

### 1st Law of Thermodynamics

- ↳ energy can be neither created nor destroyed in a chemical reaction, or
- ↳ the total sum of matter and energy in the Universe is constant, or
- ↳  $\Delta G(\text{universe}) = 0$

Note:  $\Delta G(\text{system})$  and  $\Delta G(\text{surroundings})$  are typically greater or less than zero, but the sum of the two must equal zero

## 2nd Law of Thermodynamics

- ↳ the entropy of the Universe is always increasing, or
- ↳ the energy in the Universe is always "spreading out" or dispersed, or
- ↳ the energy in the Universe is always becoming less usable, or
- ↳  $\Delta S(\text{universe}) > 0$

## Summary

For any spontaneous energy change:

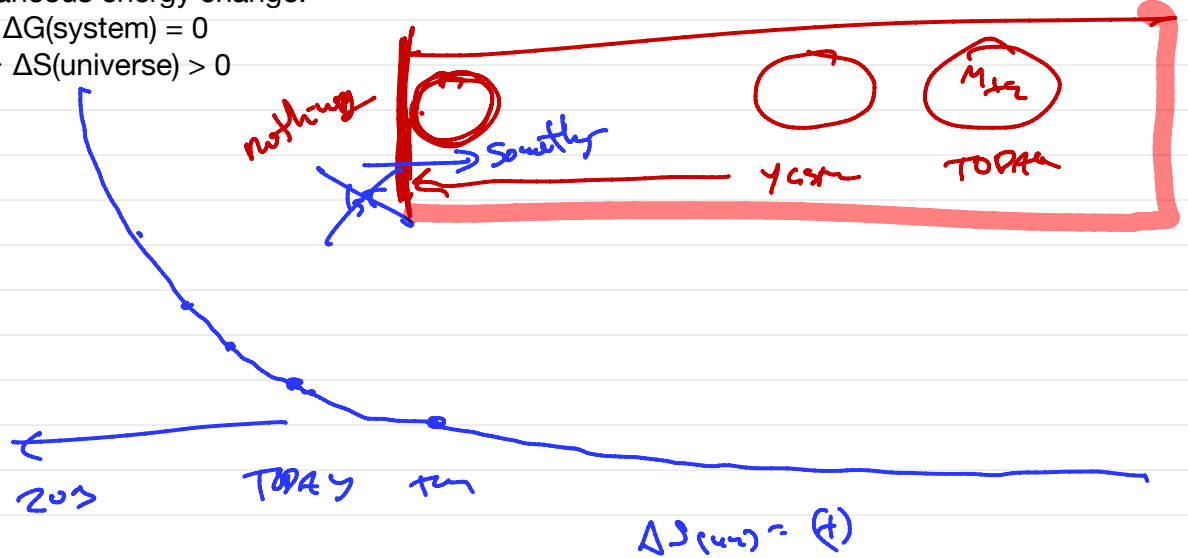
1st Law  $\rightarrow \Delta G(\text{system}) = 0$

2nd Law  $\rightarrow \Delta S(\text{universe}) > 0$

$$\Rightarrow \Delta G(\text{univ}) = 0$$

$$\text{2nd} \quad \Delta S(\text{univ}) = +$$

High  
order



## Calculating enthalpy (heat change) using Stoichiometry

The “trick” — treat heat as if it were a reactant or product

The amount of heat associated with a reaction can be easily calculated using a stoichiometry “trick” — write the enthalpy of balanced chemical equation as either a reactant or product.

### EXAMPLE 1: HEAT AS PRODUCT

Consider the reaction,  $A + B \rightarrow C$ , wherein 100 kJ of heat is GIVEN OFF.



$$\Delta H = -100 \text{ kJ}$$

↓  
exo

∴ heat = product

↳ if heat is given off, then  $\Delta H = (-)$

↳ if  $\Delta H = (-)$ , then process is EXO

↳ if EXO, then treat heat as a Product (as heat is formed)

↳ therefore, write heat/ $\Delta H$  on the right side of the equation.



## EXAMPLE 2: HEAT AS REACTANT

Consider the reaction,  $A + B \rightarrow C$ , wherein  $\Delta H = +100 \text{ kJ}$ .

↳ if  $\Delta H = (+)$ , then process is ENDO

↳ if ENDO, then treat heat as a Reactant (heat is added to maket the reaction go, it is not the result of the reaction having taken place)

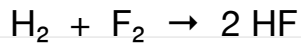
↳ therefore, write heat/ $\Delta H$  on the left side of the equation.



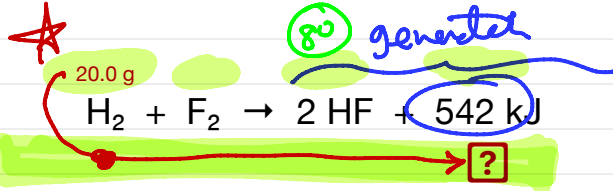
according to the following equation.

(EX) ¿For every 1 mole of H<sub>2</sub> consumed, 542 kJ of heat is generated, How much heat is produced by burning 20.0 g of H<sub>2</sub>?

- \$100  
spent \$100



exothermic → add to product-side

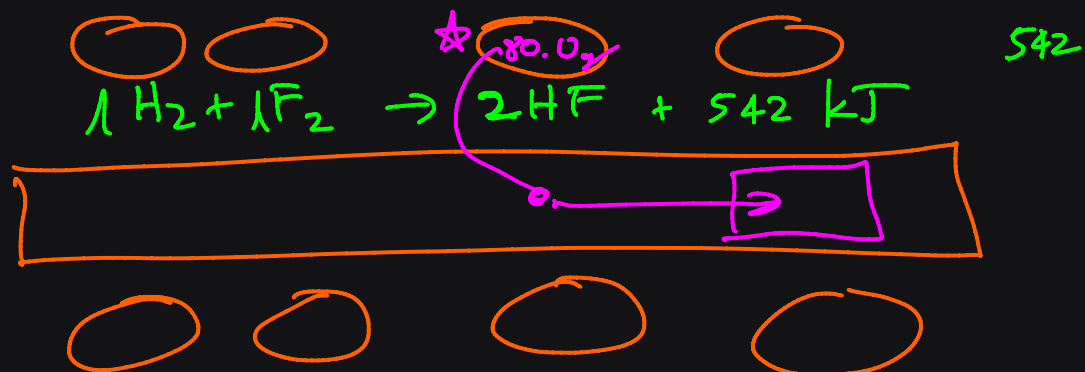


$$\Delta H = -542$$

treat "kJ" as a compound, "542" as coefficient,  
and then work stoichiometry problem as usual  
(box = star + track)

$$\frac{\boxed{\text{KJ}}}{1} = \frac{20.0 \text{ g}}{1} \times \frac{1 \text{ mol H}_2}{2 \text{ g H}_2} \times \frac{542 \text{ KJ}}{1 \text{ mol H}_2} = 5420 \text{ KJ} \left| \frac{1000 \text{ J}}{1 \text{ KJ}} \right.$$

(EX) ¿How much heat is produced during the production of 80.0 g of HF?



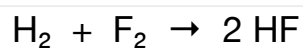
$$\frac{\cancel{542} \text{ kJ}}{1} = \frac{80.0 \cancel{\text{g HF}}}{20. \cancel{\text{g HF}}} \times \frac{1 \cancel{\text{ mol HF}}}{2 \cancel{\text{ mol HF}}} \times \frac{542 \text{ kJ}}{1}$$

19

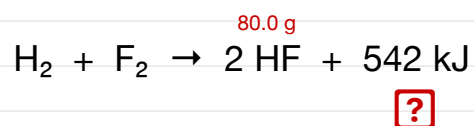
$$\frac{\text{kJ} = 15}{1 + 1} = 20.$$

(EX) ¿How much heat is produced during the production of 80.0 g of HF?

Basically, the same question as above...



heat "produced" →  
exothermic → add to product-side



← After the setup, just your basic Dim.  
Analys. problem.... Box = Star + Traintrack



treat "kJ" as a compound, "542" as coefficient,  
and then work stoichiometry problem as usual  
(box = star + track)

$$\frac{\square \text{ KJ}}{1} = \frac{80. \text{ g}}{1} \times \frac{1 \text{ mol HF}}{20. \text{ g HF}} \times \frac{542 \text{ KJ}}{2 \text{ mol HF}} = 1100 \text{ KJ}$$