Name: ____________________________________ Period:___________________

CHAPTER 17: THERMOCHEMISTRY

SPECIFIC HEAT CAPACITY

1. Units of heat:

Joule: metric unit for heat/energy

calorie: amount of energy needed to change 1 gram of water 1°C

kilocalorie (Calorie or food calorie)

2. Specific Heat Capacity how much energy it takes to change 1 gram of something's temp. by 1°C

low specific heat capacity: it takes less energy to raise the temp (ex: metal)

high specific heat capacity: it takes more energy to raise the temp (ex: water)

3. Specific Heat Capacity – think about it!

1. Which substance below would have the highest specific heat capacity? Which would have the lowest?

   a. sand, water, air water = highest, sand = lowest

   b. iron, granite, glass glass = highest, iron = lowest

2. Look at the table of specific heat capacities below.

Substance	Specific Heat Capacity
aluminum	0.897 J/g°C
iron	0.454 J/g°C
lead	0.130 J/g°C

Which substance heats up the quickest? __________ lead __________________If each of the substances above absorbs 1000J of heat, which would become the hottest? Why? Lead – it takes the least amount of energy to raise its temperature

Practice specific heat problems:

a. The specific heat capacity of aluminum is 0.897 J/g°C. Determine the amount of heat released when 10.5g of aluminum cools from 240°C to 25°C.

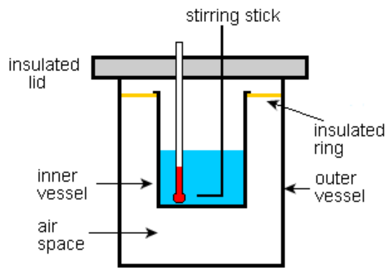
H = m x (sh) x ΔT
= 10.5g (0.897 J/g°C) (215) = 2024.98 J

b. The specific heat capacity of aluminum is 0.897 J/g°C. What mass of aluminum can be heated from 33°C to 99°C using 450J of heat?

H = m x (sh) x ΔT
450 J = m (0.897 J/g°C) (66)
7.6g = m

Calorimetry: heat gained by water = heat lost by process m(sh)(ΔT) = m(sh)(ΔT)

CALORIMETER



1. 55.0g of a "mystery metal" is heated to 93°C in a hot water bath, and is then placed in a calorimeter containing 100.0g of water at 25°C. The metal sits in the water until the temperature levels off at 29°C. At this point, both the metal and the water are at 29°C. The specific heat capacity of water is 4.18 J/g°C. Find the specific heat capacity of the metal.

	METAL	WATER IN CALORIMETER
mass	55g	100 g
specific heat capacity	?	4.18 J/g°C
$T_{initial}$	29 °C	29 °C
T_{final}	93 °C	25 °C
ΔT	64 °C	4 °C

a. Use H = m x (sh) x ΔT to find the amount of heat gained by the water.
H = 100(4.18)(4)
H = 1672 J

b. How much heat was lost by the metal?
1672 J (same amount of energy gained by the water!!)

c. Use H = m x (sh) x ΔT to find the specific heat capacity of the metal.
1672 = 55(sh)(64)
0.475 J/g°C = sh

2. 84.0g of a metal are heated to 99ºC in a hot water bath, and then placed in a coffee cup calorimeter containing 60.0g of water at 32ºC. The final temperature in the calorimeter is 42ºC. What is the specific heat of the metal?

	METAL	WATER IN CALORIMETER
mass	84.0 g	60.0g
specific heat capacity	?	4.18 J/g°C
$T_{initial}$	99ºC	32ºC
T_{final}	42ºC	42ºC
ΔT	57 ºC	10 ºC

m(sh)(ΔT) = m(sh)(ΔT)
84.0g(sh)(57 ºC) = 60.0g(4.18 J/g°C)(10 ºC)
sh = 0.52 J/g°C

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3. The specific heat of aluminum is 0.90 J/gºC. 30.0g of aluminum are heated in a hot water bath to 98ºC, and added to a calorimeter containing water at 25ºC. The final temperature in the calorimeter is 32ºC. What mass of water was in the calorimeter?

	METAL	WATER IN CALORIMETER
mass	30.0g	??
specific heat capacity	0.90 J/gºC	4.18 J/g°C
$T_{initial}$	98ºC	25ºC
T_{final}	32ºC	32ºC
ΔT	66 ºC	7 ºC

m(sh)(ΔT) = m(sh)(ΔT)
30.0g(0.90 J/gºC)(66 ºC) = m(4.18 J/g°C)(7 ºC)
sh = 60.9g

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4. 124.0g of aluminum (sh = 0.90 J/g°C) are heated to 99ºC in a hot water bath, and then placed in a coffee cup calorimeter containing 65.0g of water. The final temperature in the calorimeter is 43ºC. What was the initial temperature of the water?

	METAL	WATER IN CALORIMETER
mass	124.0g	65.0g
specific heat capacity	0.90 J/g°C	4.18 J/g°C
$T_{initial}$	99ºC	
T_{final}	43ºC	43ºC
ΔT	56 ºC	X = 23.0 ºC

m(sh)(ΔT) = m(sh)(ΔT)
124.0g(0.90 J/gºC)(56 ºC) = 65.0g(4.18 J/g°C)(ΔT)
ΔT = 23.0 ºC

43 ºC + 23.0 ºC = 66.0 ºC

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*****5. 125.0g of copper (sh ≈ 0.40 J/g°C) are heated to 97ºC in a hot water bath, and then placed in a coffee cup calorimeter containing 42.0g of water at 23°C. Find the final temperature in the calorimeter.**

	METAL	WATER IN CALORIMETER
mass	125.0g	42.0g
specific heat capacity	0.40 J/g°C	4.18 J/g°C
$T_{initial}$	97ºC	23°C
T_{final}	X	X
ΔT	97.0°C -Tf	Tf-23.0°C

m(sh)(ΔT) = m(sh)(ΔT)
125.0g (0.40 J/g°C) (97.0°C -Tf) = 42.0g (4.18 J/g°C) (Tf-23.0°C)
4850 – 50Tf = 175.6Tf – 4037.9
8887.9 = 225.6Tf
39.4 ºC = Tf

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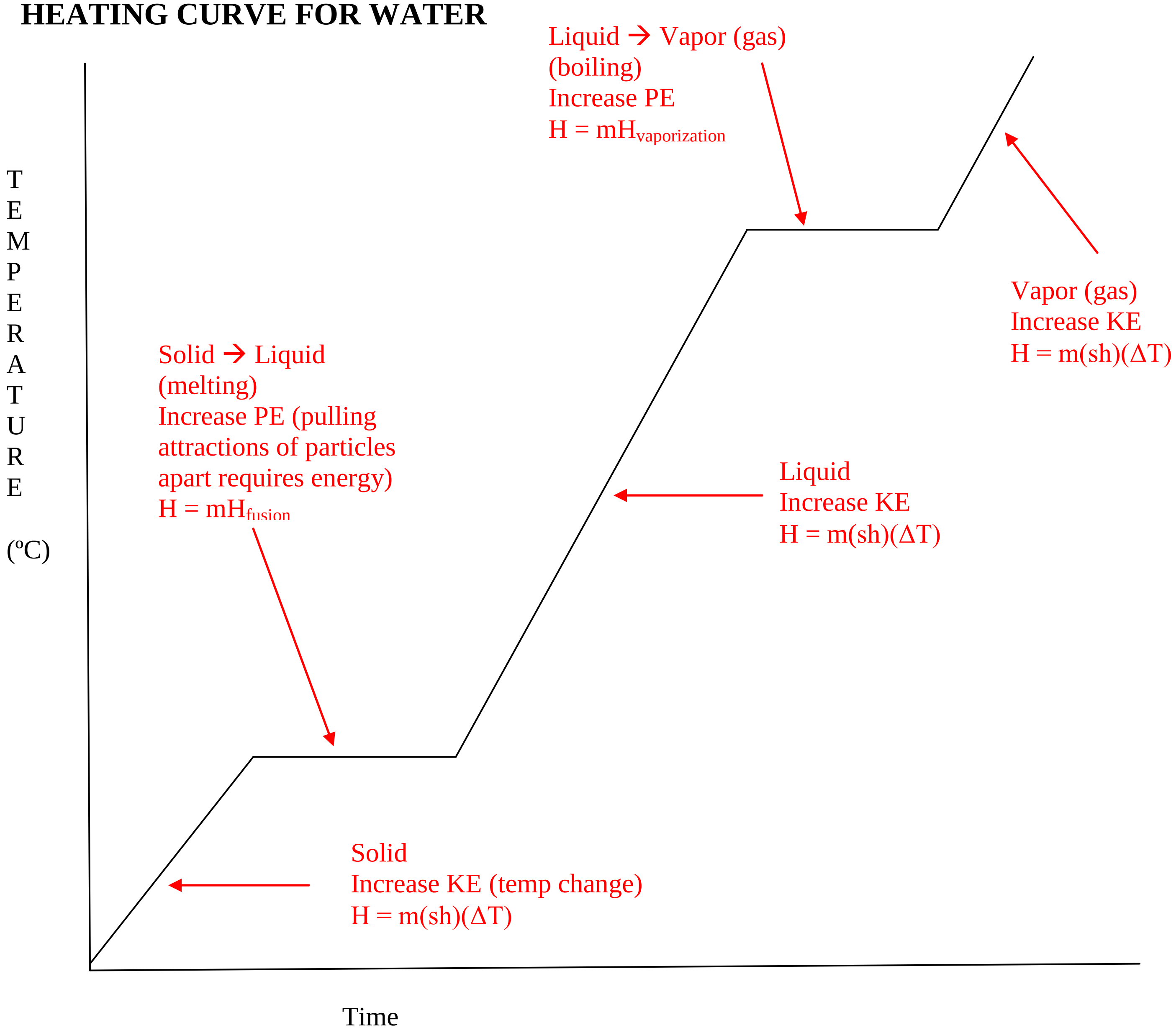
HEATING CURVE FOR WATER



Vocabulary:

heat of fusion: **energy necessary to melt one gram of a substance at its melting point, energy given off when one gram of a substance freezes at its freezing point**

heat of vaporization: **energy needed to boil one gram of a substance at its boiling point, energy released if a gram of a substance condenses**

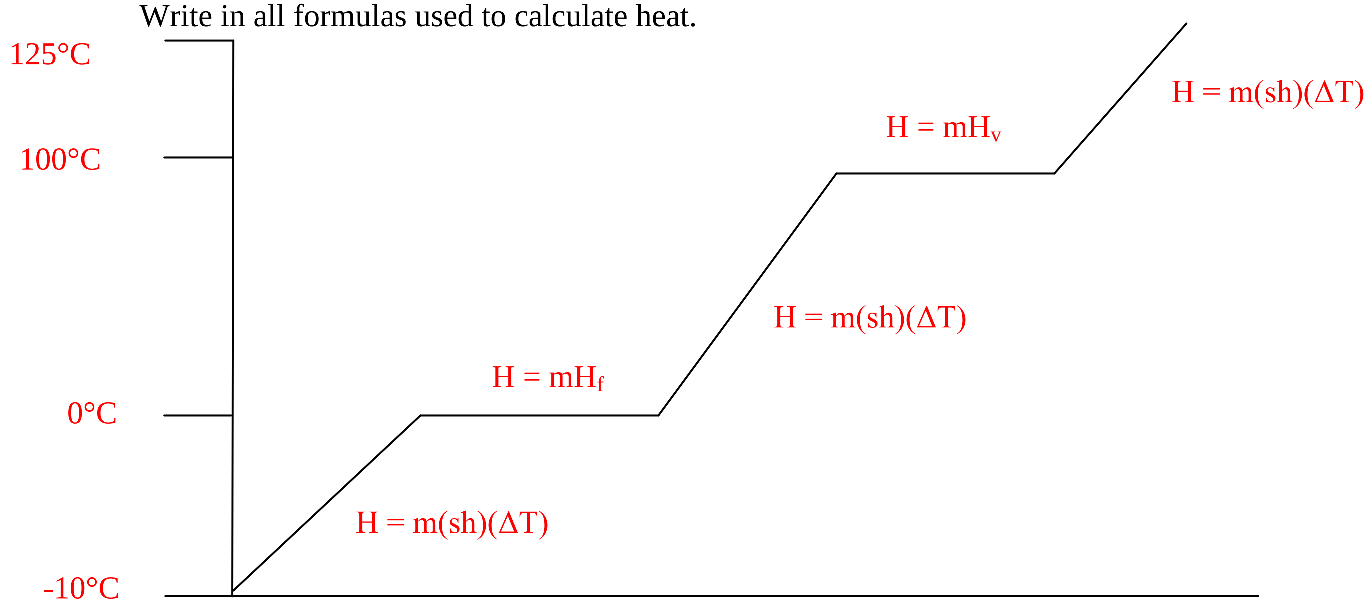
kinetic energy: **energy of motion (temperature)**

potential energy: **energy of position (stored energy) – energy you have because you "are"**

The data below are for water (H_2O)

Melting point	Boiling Point	Heat of Fusion	Heat of Vaporization	SH Capacity (solid)	SH Capacity (liquid)	SH Capacity (vapor)
0.0°C	100.0°C	334 J/g	2260 J/g	2.05 J/g°C	4.18 J/g°C	1.90 J/g°C

1. Draw a heating curve for substance water, going from -20°C to 125°C on the axis below. Write in all formulas used to calculate heat.



b. Determine the amount of heat necessary to convert 15.0g of water at 100.0°C to steam at 100.0°C.

$H = mH_v$
$H = 15(2260)$

H = 33,900 J

ΔH = __+___

c. Determine the amount of heat necessary to convert 15.0g of ice at 0.0°C to liquid at 0.0° C.

$H = mH_f$
$H = 15(334)$

H = 5010 J

ΔH = __+___

d. Determine the amount of heat given off when 25g of water cools from 56.2°C to 33.5°C.

$H = m(sh)(\Delta T)$
$H = 25g(4.18\ J/g°C)(22.7°C)$

H = 2372.15 J

ΔH = _-____

e. Determine the amount of heat needed to raise the temperature of 19.6g of X from -15°C to -2.0°C.

$H = m(sh)(\Delta T)$
$H = 19.6g(2.05\ J/g°C)(13°C)$

H = 522.34 J

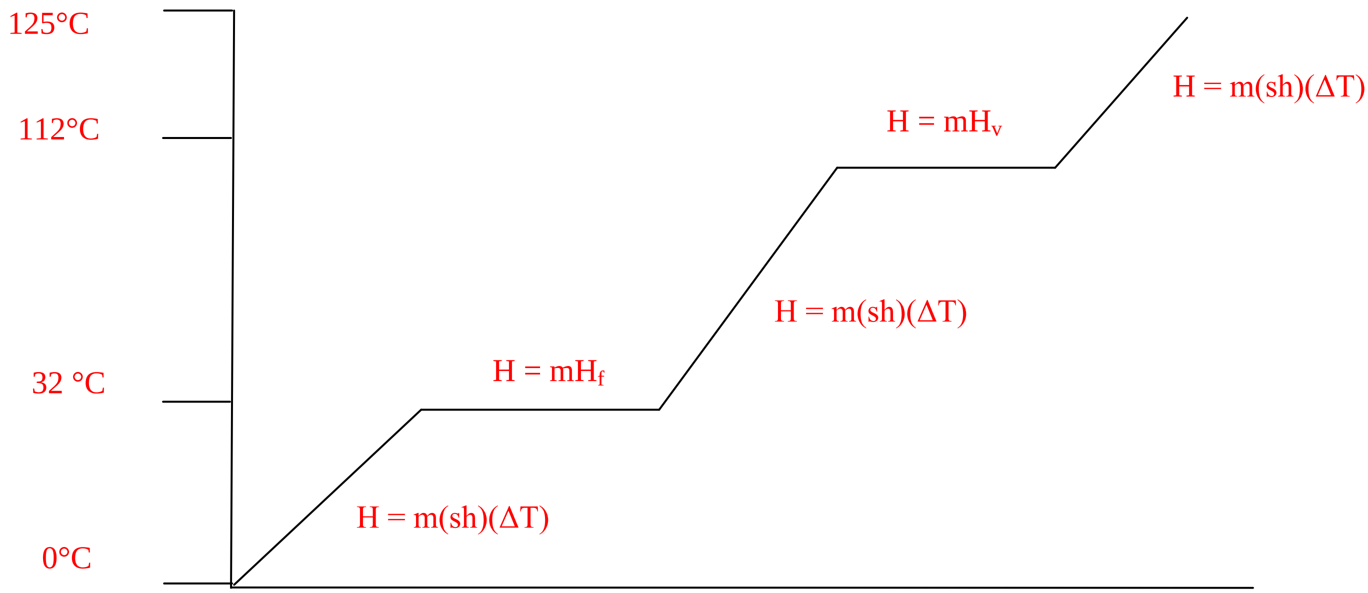
ΔH = ___-__

MORE HEAT PRACTICE

The data below refers to an unknown substance, X.

Melting point	Boiling Point	Heat of Fusion	Heat of Vaporization	SH Capacity (solid)	SH (liquid)	SH (vapor)
32.0°C	112.0°C	425 kJ/g	695 kJ/g	2.3 J/g°C	5.9 J/g°C	1.1 J/g°C

2. Draw a heating curve for substance X, going from 15°C to 125°C on the axis below. Write in all formulas used to calculate heat.



b. Determine the amount of heat released when 15.0g of gaseous X at 112.0°C changes to liquid at 112.0°C.

$H = mH_v$
$H = 15g\ (695\ kJ/g)$

H = 10425 J

ΔH = _-_____

c. Determine the amount of heat necessary to convert 15.0g of solid X at 32.0°C to liquid at 32.0° C.

$H = mH_f$
$H = 15g\ (425\ kJ/g)$

H = 6375 J

ΔH = __+___

d. Determine the mass of X that can be heated from 38°C to 102°C using 4500J of heat.

$H = m(sh)(\Delta T)$
$4500\ J = m\ (5.9\ J/g°C)\ (64\ °C)$

m = 11.9 g

e. Find the final temperature if 150g of X, at 19°C, receives 650J of heat.

$H = m(sh)(\Delta T)$
$650\ J = 150\ (2.3\ J/g°C)\ (T_f - 19\ °C)$

T_f = 20.88 °C

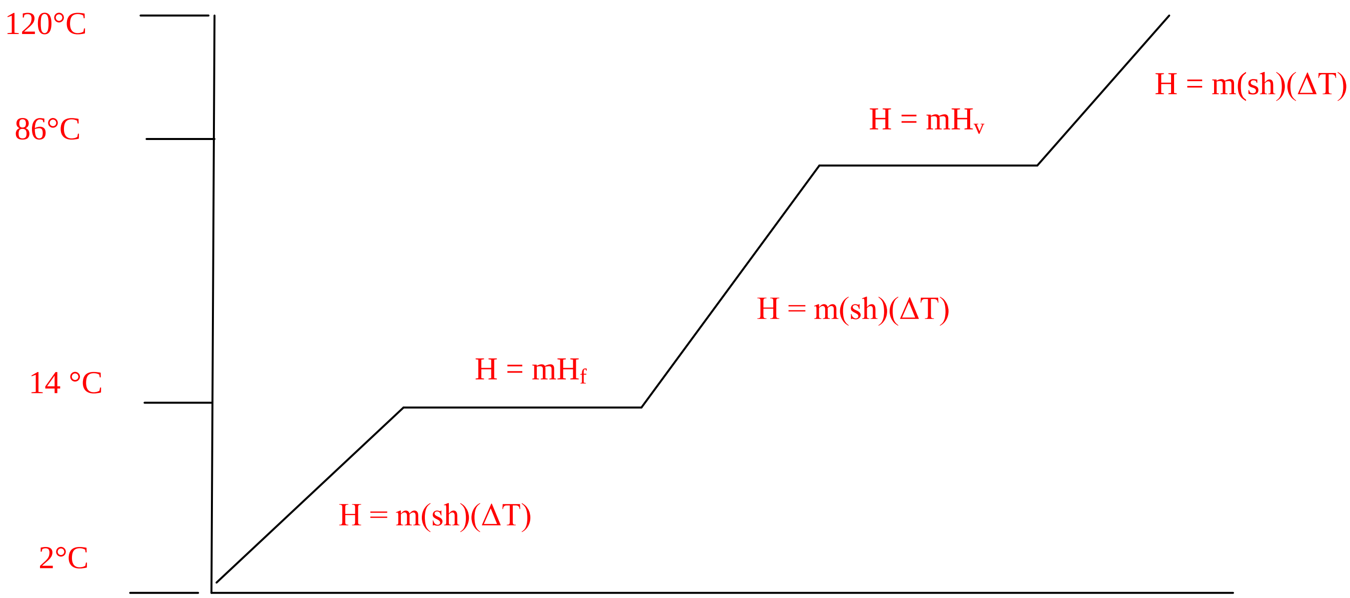
ΔH = __+___

MULTI-STEP HEAT PROBLEMS

The data below refers to an unknown substance, X.

Melting point	Boiling Point	Heat of Fusion	Heat of Vaporization	SH Capacity (solid)	SH (liquid)	SH (vapor)
14.0°C	86.0°C	150 J/g	550 J/g	4.3 J/g°C	5.2 J/g°C	1.1 J/g°C

1. Draw a heating curve for substance X, going from 2°C to 120°C on the axis below. Write in all formulas used to calculate heat.



b. Determine the amount of heat necessary to change 30g of X from solid at 4°C to liquid at 14°C.

STEP 1: $H = m(sh)(\Delta T)$
$= 30g\ (4.3\ J/g°C)\ (14-4)$
$= 1290\ J$

STEP 2: $H = mH_f$
$= (30g)\ (150\ J/g)$
$= 4500\ J$

TOTAL: 1290 J + 4500 J = 5790 J

c. How much heat will be released when 25g of X cools from 110°C to 50°C?

STEP 1: $H = m(sh)(\Delta T)$
$= (25g)\ (1.1\ J/g°C)\ (110-86)$
$= 660\ J$

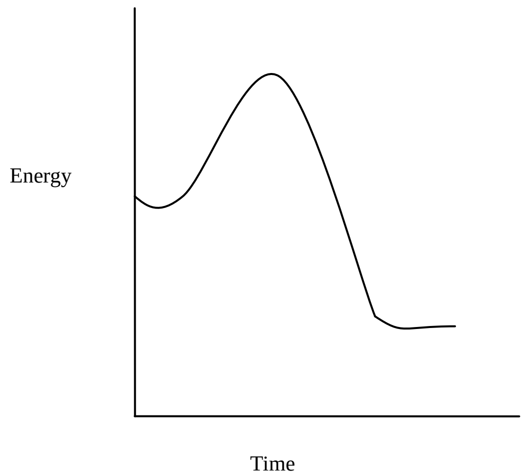
STEP 2: $H = mH_v$
$= 25g\ (550\ J/g)$
$= 137350\ J$

STEP 3: $H = m(sh)(\Delta T)$
$= 25\ (5.2\ J/g°C)\ (86-50)$
$= 4680\ J$

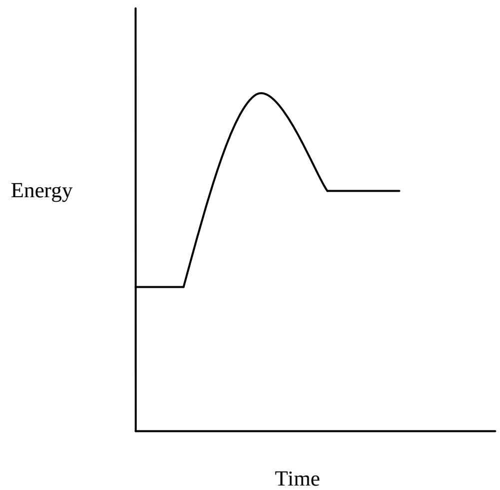
TOTAL: 660 + 137350 + 4680 = 142690 J

ENTHALPY OF CHEMICAL REACTIONS

Exothermic Reaction: ΔH is –
(heat is a "product" – right side of rxn)



Endothermic Reaction: ΔH = +
(heat is a "reactant" – left side of rxn)



ENERGY AND STOICHIOMETRY

1. Consider the following reaction:

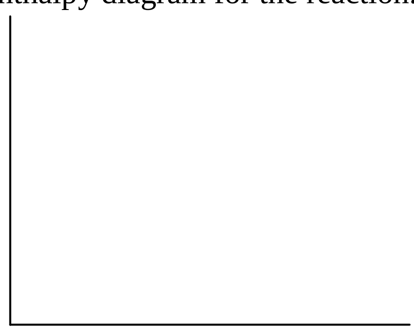
$$2\ S\ +\ 3\ O_2 \rightarrow 2\ SO_3\ +\ 791.4kJ$$

a. Does the reaction absorb or release heat? ________________

b. Is the reaction endothermic or exothermic? ________________

c. What is the value of ΔH? ________________

d. Draw an enthalpy diagram for the reaction.



e. How much heat will be given off if 10.0g of sulfur burn in excess oxygen?

f. What mass of sulfur must burn in excess oxygen in order to release 250kJ of heat?

2. Consider the following reaction:

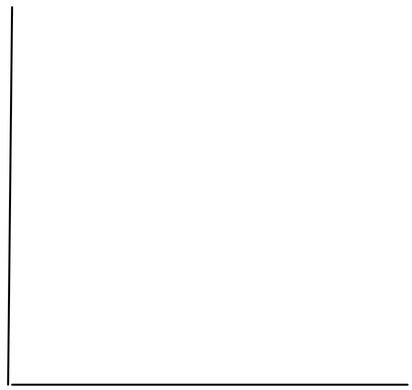
$$H_2 \;+\; Br_2 \rightarrow 2\,HBr \qquad\qquad \Delta H = 72.80kJ$$

a. Does the reaction absorb or release heat? ______________________

b. Is the reaction endothermic or exothermic? _____________________

c. On which side of the reaction would heat appear? ________________

d. Draw an enthalpy diagram for the reaction.



e. How much heat is required to form 125g of HBr?

f. What mass of HBr can be formed if 525kJ of heat are absorbed by the reaction?

THERMODYNAMICS: ENERGY & ENTROPY IN CHEMICAL REACTIONS

Enthalpy of a Reaction (ΔH): change in heat ($H_{products} - H_{reactants}$)

ΔH = "+" = endothermic = H absorbed

ΔH = "-" = exothermic = H released

Entropy: (ΔS) – change in order/how disorganized a substance is

ΔS = "+" = system is more disorganized

ΔS = "-" = system is less disorganized

Which process(es) below would have a positive change in entropy (ΔS)?

a. cleaning your room
   - ΔS
b. a bottle breaking
   + ΔS
c. leaves falling off of a tree
   + ΔS
d. water freezing
   - ΔS

States of matter, solutions and entropy:

Solid Liquid Aqueous (solution) Gas

Determine the signs of ΔH and ΔS for:

a. ice melting	ΔH = ___+______	ΔS = __+_______
b. steam condensing	ΔH = ___-______	ΔS = __-______
c. sublimation	ΔH = ___+______	ΔS = ___+______

Which reactions below would have a positive change in entropy (ΔS)?

a. $Al_2(CO_3)_3(s) \rightarrow Al_2O_3(s) + 3CO_2(g)$

Positive ΔS

b. $KCl(aq) + Br_2(l) \rightarrow KBr(aq) + Cl_2(g)$

Positive ΔS

c. $FeCl_2(aq) + H_2SO_4(aq) \rightarrow FeSO_4(s) + HCl(aq)$

Negative ΔS

How can you tell if a reaction will be *spontaneous*? (able to occur?)

ΔS change in entropy	ΔH change in heat	Spontaneous???
+ more disorder	- exothermic releases heat	Yes – it can occur at any temp.
+ more disorder	+ endothermic requires heat	Spontaneous at high temperature
- less disorder/more order	- exothermic releases heat	Spontaneous at low temperature
- less disorder/more order	+ endothermic requires heat	No – it cannot occur

1. **Determine whether each reaction will have a positive or negative change in entropy (ΔS).**
2. **Label each reaction as endothermic or exothermic.**
3. **Determine which reaction(s) below will naturally occur (be spontaneous)? Which will require added energy? Which will only occur naturally at low temperatures?**

a. $Na_2CO_{3\,(s)} \rightarrow Na_2O_{(s)} + CO_{2\,(g)}$ ΔS = __+___ $\Delta H = 321kJ$

Spontaneous? ___@ high temp__

b. $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$ ΔS = __+__ $\Delta H = -457.6kJ$

Spontaneous? ___always______

c. $N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$ ΔS = _-____ $\Delta H = -91.8kJ$

Spontaneous? _@ low temp______

d. $2KCl_{(s)} + 3O_{2\,(g)} \rightarrow 2KClO_{3\,(s)}$ ΔS = __-___ $\Delta H = 90.0kJ$

Spontaneous? __no___________