ENERGY

- thermodynamics: the study of energy transfer

Conservation of energy: Energy may change form, but the overall amount of energy remains constant. "first law of thermodynamics"

- ... but what IS energy?
  - energy is the ability to do "work"

motion of matter

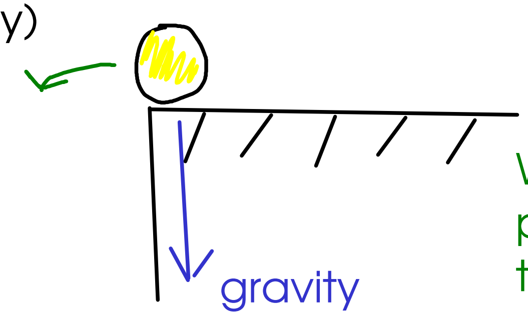
Kinds of energy?

- Kinetic energy: energy of matter in motion $E_K = \frac{1}{2} m v^2$

mass

velocity

- Potential energy: energy of matter that is being acted on by a field of force (like gravity)



When the ball falls, its potential energy is converted to kinetic!

- What sort of energy concerns chemists? Energy that is absorbed or released during chemical reactions.

- Energy can be stored in chemicals ... molecules and atoms.

INTERNAL ENERGY: "U"

related to the kinetic and potential energy of atoms, molecules, and their component parts.

- We measure energy transfer ... which is called HEAT. (HEAT is the flow of energy from an area of higher temperature to an area of lower temperature)

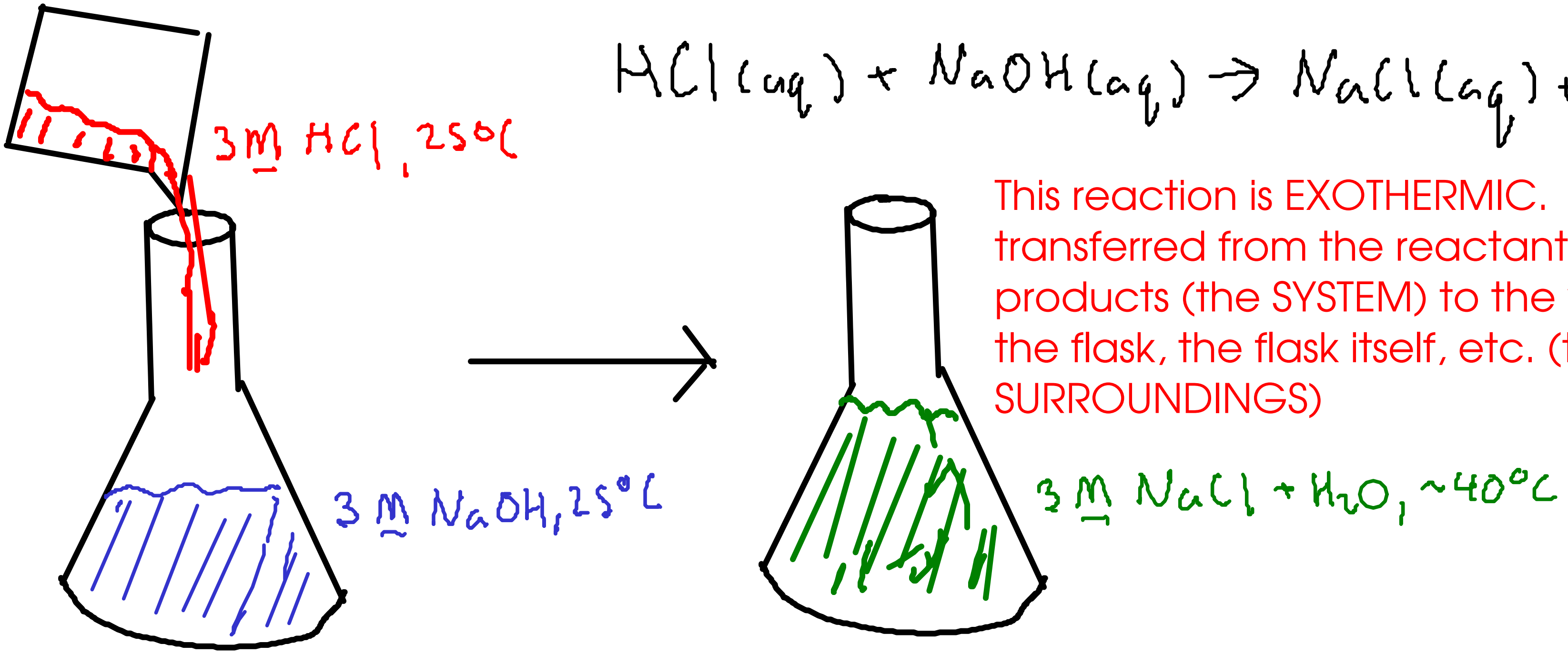
Q : heat

SYSTEM: the object or material under study

SURROUNDINGS: everything else

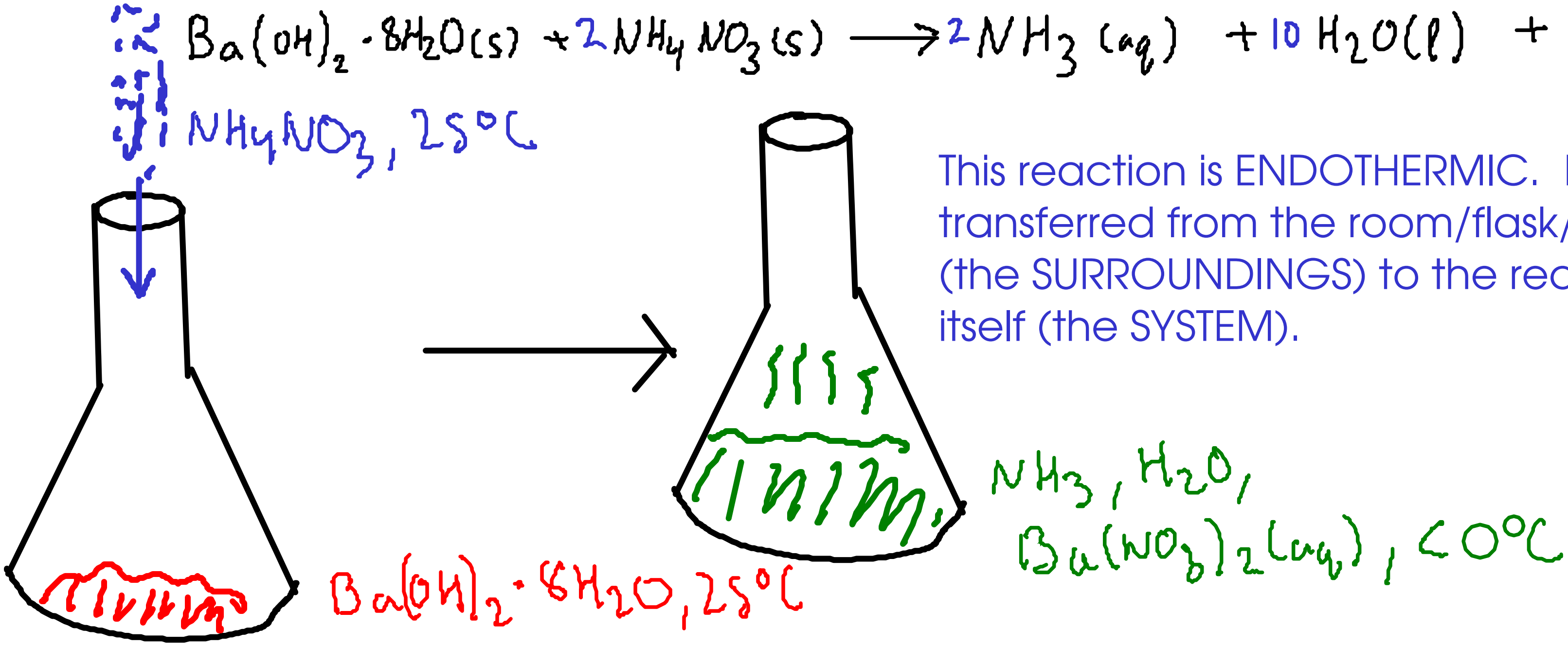
Type of process	Energy is ...	Sign of Q	Temp of SURROUNDINGS ...
ENDOTHERMIC	transferred from SURROUNDINGS to SYSTEM	+	decreases
EXOTHERMIC	transferred from SYSTEM to SURROUNDINGS	−	increases

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$



This reaction is EXOTHERMIC. Energy is transferred from the reactants and products (the SYSTEM) to the water in the flask, the flask itself, etc. (the SURROUNDINGS)

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4NO_3(s) \longrightarrow 2NH_3(aq) + 10H_2O(l) + Ba(NO_3)_2(aq)$$

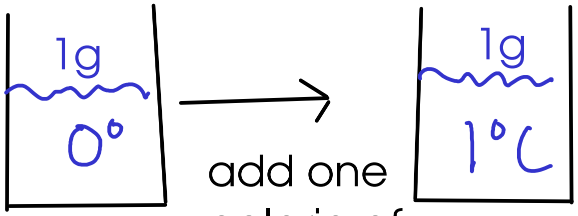


This reaction is ENDOTHERMIC. Energy is transferred from the room/flask/etc. (the SURROUNDINGS) to the reaction itself (the SYSTEM).

ENERGY UNITS

- calorie (cal): the amount of energy required to change the temperature of one gram of water by one degree Celsius (or Kelvin)

1g ≈ 1mL for water



add one
calorie of
energy

- Calories in food? The "Calorie" that is given on American food labels is actually the kilocalorie (kcal)

- Joule (J): SI unit for energy. It's defined based on the equation for kinetic energy.

$$1\,J = 1\,\frac{kg\,m^2}{s^2}$$

, from

$$E_K = \frac{1}{2} m v^2$$

kinetic energy

mass

velocity

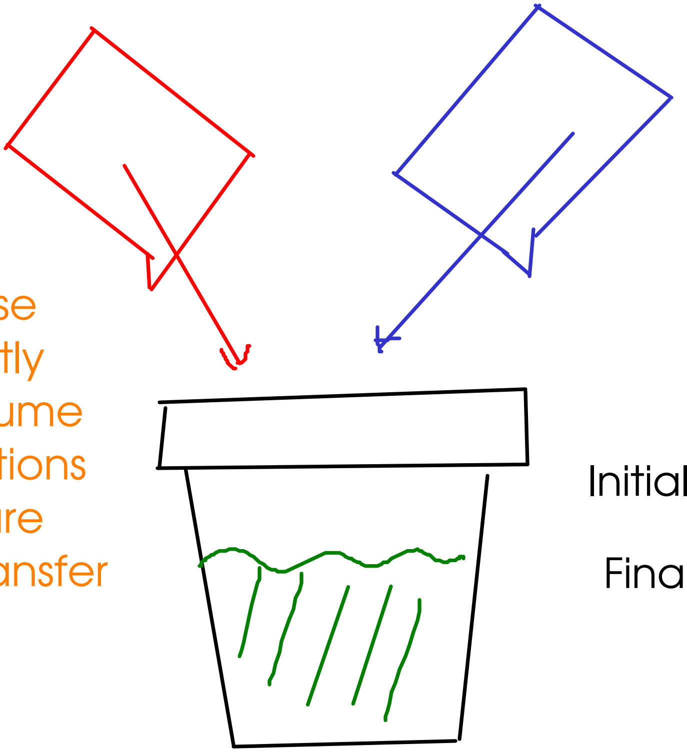
$\boxed{4.184\ J = 1\ cal}$

- the Joule is a small unit. For most reactions at lab scale, we'll use kilojoules (kJ).

CALORIMETRY - measurement of heat

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

50. mL
2.0 M HCl



50. mL
2.0 M NaOH

The acid and base solutions are mostly water ... we'll assume that the two solutions are essentially pure water for heat transfer purposes...

Initial temperature of solutions: 22.1° C

Final temperature of mixture: 35.6° C

Assuming that no heat is lost from the water to the surrounding air,

$$\underbrace{Q_r}_{\text{reaction}} + \underbrace{Q_w}_{\text{water}} = 0$$

Conservation of energy. The terms add to zero because they have opposite signs.

... if we knew something about the WATER, we could use that to find the heat of the REACTION!

SPECIFIC HEAT

- a measured quantity. The amount of energy required to change the temperature of one gram of a particular substance by one degree Celsius.

- Specific heat information for common substances is readily available. For water,

$$4.184 \frac{J}{g\,°C} \quad \text{or} \quad 1.000 \frac{cal}{g\,°C}$$

$$Q = m \times S \times \Delta T$$

m = mass
s = specific heat
Δ T = Tfinal - Tinitial

This is ALWAYS final temp minus initial temp!

- For objects, like reaction vessels, you might know the HEAT CAPACITY, which is the amount of energy required to change the temperature of an object by one degree Celsius

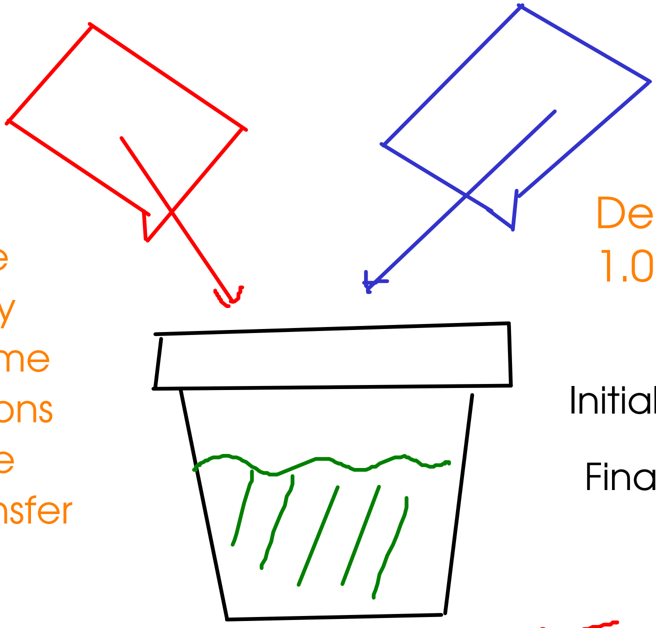
Units: $J/°C$ or $cal/°C$

$$Q = C \times \Delta T$$

c = heat capacity

$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

50. mL
2.0 M HCl



50. mL
2.0 M NaOH

Density of aqueous solutions is about 1.0 g/mL

The acid and bsae solutions are mostly water ... we'll assume that the two solutions are essentially pure water for heat transfer purposes...

Initial temperature of solutions: 22.1° C

Final temperature of mixture: 35.6° C

50+50

$Q_r + Q_w = 0$

$Q_w = m_w \times s_w \times \Delta T_w = (100.\,g)\left(4.184 \frac{J}{g°C}\right)(35.6°C - 22.1°C)$

$Q_w = 5648.4\,J$

$Q_r + 5648.4\,J = 0\,;\ Q_r = -5648.4\,J$

To report the energy change in this reaction to others, we should express it in terms of heat transfer per mole of something. A different amount of reactant would have a different Q

$0.050\,L \times \frac{2.0\text{ mol HCl}}{L} = 0.10\text{ mol HCl}$

"Heat of reaction"

$Q_{rxn} = \frac{Q_r}{\text{moles HCl}} = \frac{-5648.4\,J}{0.10\text{ mol HCl}} = -56000 \frac{J}{mol} = \boxed{-56 \frac{kJ}{\text{mol HCl}}}$

One problem ...

PATH. The amount of energy required for a process depends on how the process is carried out.

Example: Driving from Florence to Columbia. How much energy is required? (gas)

Jeep Cherokee (15 mpg) vs Toyota Prius (43 mpg). The Jeep will use much more fuel than the Prius even though they start and end at the same place. So the fuel usage is what we call a PATH FUNCTION, while the location is a STATE FUNCTION.

- so the heat of reaction depends on how the reaction is done.
  - we need (for reporting) some kind of standard condition. At constant pressure, we can define a state function called ENTHALPY (H)

$$H = U + PV$$

$$\Delta H = Q_{\text{constant pressure}}$$

... we record the "enthalpy change of reaction" in our data books.

$$\Delta H_r$$

SINCE the enthalpy change does NOT depend on path, this means that we can use standard values for enthalpy to predict the heat change in reactions that we have not tested in a calorimeter.

THERMOCHEMICAL EQUATIONS

- are like regular chemical equations, except that phase labels are REQUIRED and the enthalpy for the reaction is given.

$$CH_3COCH_3(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l)\ ;\ \Delta H = -1800\ kJ$$

- Why are phase labels required? Because phase changes either absorb or release energy.

$\Delta H = -1800\ kJ$ ... what does this mean?

$1\ mol\ CH_3COCH_3 = -1800\ kJ$

$4\ mol\ O_2 = -1800\ kJ$

$3\ mol\ CO_2 = -1800\ kJ$

$3\ mol\ H_2O = -1800\ kJ$

We treat the enthalpy change as if it's another product of the reaction!

A few more terms related to enthalpy:

- Enthalpy of vaporization / heat of vaporization: The enthalpy change on vaporizing one mole of a substance. (from liquid to vapor)

- Enthalpy of fusion / heat of fusion: The enthalpy change when a mole of liquid changes to the solid state.

$$CH_3COCH_3(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l)\ ;\ \Delta H = -1800\ kJ$$

What would be the enthapy change when 25 g of water are produced by the reaction?

1- Convert 25.0 g water to moles using FORMULA WEIGHT.
2 - Convert moles water to ENTHALPY CHANGE using THERMOCHEMICAL EQUATION

$18.016\ g\ H_2O = mol\ H_2O$

$3\ mol\ H_2O = -1800\ kJ$

$$25\ g\ H_2O \times \underset{\textcircled{1}}{\frac{mol\ H_2O}{18.016\ g\ H_2O}} \times \underset{\textcircled{2}}{\frac{-1800\ kJ}{3\ mol\ H_2O}} = \boxed{-830\ kJ}$$

This also equals Q (heat), as long as pressure is constant.

This is an EXOTHERMIC reaction. Not surprising, since it's a COMBUSTION reaction!

FORMATION REACTIONS

- A reaction that forms exactly one mole of the specified substance from its elements at their STANDARD STATE at 25C and 1 atm pressure.

$CO_2(g)$: $C(s, graphite) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = \underline{-393.5\ kJ}$

heat of formation of carbon dioxide ΔH_f° or ΔH_f
...also called "enthalpy of formation"

$CO(g)$: $C(s, graphite) + \underline{\underline{\frac{1}{2}}}O_2(g) \rightarrow CO(g)$; $\Delta H = -110.5\ kJ$

you may see fractional coefficients in these formation reactions, because you MUST form exactly one mole of the product!

- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

$$\Delta H^\circ_{f, O_2(g)} = 0\ kJ/mol$$

- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

Let's say we would like to find the enthalpy of reaction for this equation:

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g);\ \Delta H = ?$$

From A-8, text ↙

Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.

$$C_2H_2:\quad 2C(s) + H_2(g) \longrightarrow C_2H_2(g) \qquad \Delta H = 226.7$$

$$CO_2:\quad C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5$$

$$H_2O:\quad H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(g) \qquad \Delta H = -241.8$$

$$C_2H_2(g) \longrightarrow \cancel{2C(s)} + \cancel{H_2(g)} \qquad \Delta H = -226.7$$

$$C_2H_2(g) \longrightarrow \cancel{2C(s)} + \cancel{H_2(g)} \qquad \Delta H = -226.7$$

$$\cancel{C(s)} + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5$$

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$$\cancel{H_2(g)} + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(g) \qquad \Delta H = -241.8$$

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l)$$

$$\Delta H = 2(-226.7) + 4(-393.5) + 2(-241.8) = \boxed{-2511\ \text{kJ}}$$

Hess' Law using enthalpy of formation:

$$\Delta H = \sum \Delta H_{f,\text{products}} - \sum \Delta H_{f,\text{reactants}}$$

$$2\overset{226.7}{C_2H_2}(g) + 5\overset{0}{O_2}(g) \longrightarrow 4\overset{-393.5}{CO_2}(g) + 2\overset{-241.8}{H_2O}(g)\ ;\ \Delta H = ?$$

$$\Delta H = [4(-393.5) + 2(-241.8)] - [2(226.7) + 5(0)]$$

$$= \boxed{-2511\ kJ}$$

* Remember:

- Multiply each enthalpy by its stoichiometric coefficient from the reaction
- Enthalpy of formation of an element at its standard state is zero
- Watch phase labels. You will usually find SEVERAL enthalpies of formation for a given substance in different phases!
- For ionic substances in solution, remember that they exist as free ions, so look up the aqueous IONS!

See Appendix C in the textbook for enthalpy of formation data:
p A-8 to A-11

Example problems:

FORMULA WEIGHTS in g/mol
in purple

$$2\overset{2.016}{H_2}(g) + \overset{32.00}{O_2}(g) \longrightarrow 2\overset{18.02}{H_2O}(g)\;;\; \Delta H = -484\ \text{kJ}$$

Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

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1 - Convert 1.00 kg hydrogen gas to moles. Use FORMULA WEIGHT. (and kg -g conversion)
2 - Convert moles hydrogen gas to enthalpy change using THERMOCHEMICAL EQUATION

$2.016\ g\ H_2 = \text{mol}\ H_2$ | $2\ \text{mol}\ H_2 = -484\ \text{kJ}$ | $\text{kg} = 10^3\ g$

$$1.00\ \text{kg}\ H_2 \times \frac{10^3\ g}{\text{kg}} \times \underset{\textcircled{1}}{\frac{\text{mol}\ H_2}{2.016\ g\ H_2}} \times \underset{\textcircled{2}}{\frac{-484\ \text{kJ}}{2\ \text{mol}\ H_2}} = \boxed{\begin{array}{c}-12\bar{0}000\ \text{kJ} \\ \text{per kg}\ H_2\end{array}}$$

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$; $\Delta H = -906$ kJ

What is the enthalpy change when 150. L of nitrogen monoxide are formed by this reaction at 25.0 C and 1.50 atm pressure?

1 - Convert 150. L NO to moles using IDEAL GAS EQUATION.
2 - Convert moles NO to enthakpy change using THERMOCHEMICAK EQUATION

① $PV = nRT$

$n = \frac{PV}{RT}$

$P = 1.50 \text{ atm}$ $V = 150.\text{ L}$ $T = 25.0°C = 298.2K$

$R = 0.08206 \frac{L\text{-}atm}{mol \cdot K}$

$$n_{NO} = \frac{(1.50 \text{ atm})(150.\text{ L})}{\left(0.08206 \frac{L\text{-}atm}{mol \cdot K}\right)(298.2K)} = 9.194822849 \text{ mol } NO$$

② 4 mol NO = −906 kJ

$$9.194822849 \text{ mol } NO \times \frac{-906 \text{ kJ}}{4 \text{ mol } NO} = \boxed{-2080 \text{ kJ}}$$

Heat of formation / enthalpy of formation!

34.086 g/mol

-20.50   0   -285.8   -296.8   ] ΔH_f^o, kJ/mol (Appendix C)

$$2H_2S(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2SO_2(g)$$

What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

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1 - Calculate the enthalpy change for the reaction AS WRITTEN with HESS'S LAW

2 - Convert 25.0 grams hydrogen sulfide to moles using FORMULA WEIGHT.

3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL EQUATION.

$$\Delta H = \sum \Delta H_f^o\text{, products} - \sum \Delta H_f^o\text{, reactants}$$

$$= [2(-285.8) + 2(-246.8)] - [2(-20.50) + 3(0)]$$

$$= -1124.2\ kJ$$ ①

So the THERMOCHEMICAL EQUATION is ...

$$2H_2S(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2SO_2(g);\ \Delta H = -1124.2\ kJ$$

34.086 g H_2S = mol H_2S | 2 mol H_2S = -1124.2 kJ

$$25.0\ g\ H_2S \times \frac{mol\ H_2S}{34.086\ g\ H_2S} \times \frac{-1124.2\ kJ}{2\ mol\ H_2S} = \boxed{-412\ kJ}$$

② ③

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) ; \Delta H = -2043 \text{ kJ}$$

Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above. Propane: C_3H_8

1 - Convert the energy requirement to moles propane using THERMOCHEMICAL EQUATION
2 - Convert moles propane to volume using IDEAL GAS EQUATION

$$\text{mol } C_3H_8 = -2043 \text{ kJ}$$

Snce the reaction is the SYSTEM here, the energy requirement has a NEGATIVE sign (it's LOSING, or providing) the energy...

$$-565 \text{ kJ} \times \frac{\text{mol } C_3H_8}{-2043 \text{ kJ}} = 0.2765540871 \text{ mol } C_3H_8$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$P = 1.08 \text{ atm} \quad R = 0.08206 \frac{\text{L-atm}}{\text{mol-K}}$$

$$T = 25.0^\circ C = 298.2 \text{ K}$$

$$n = 0.2765540871 \text{ mol } C_3H_8$$

$$V = \frac{(0.2765540871 \text{ mol } C_3H_8)(0.08206 \frac{\text{L-atm}}{\text{mol-K}})(298.2 \text{ K})}{(1.08 \text{ atm})}$$

$$= \boxed{6.27 \text{ L } C_3H_8 \text{ @ } \begin{matrix} 25.0^\circ C \\ 1.08 \text{ atm} \end{matrix}}$$