

GENERAL CHEMISTRY 2

PREPARED BY:

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Special Science Teacher I

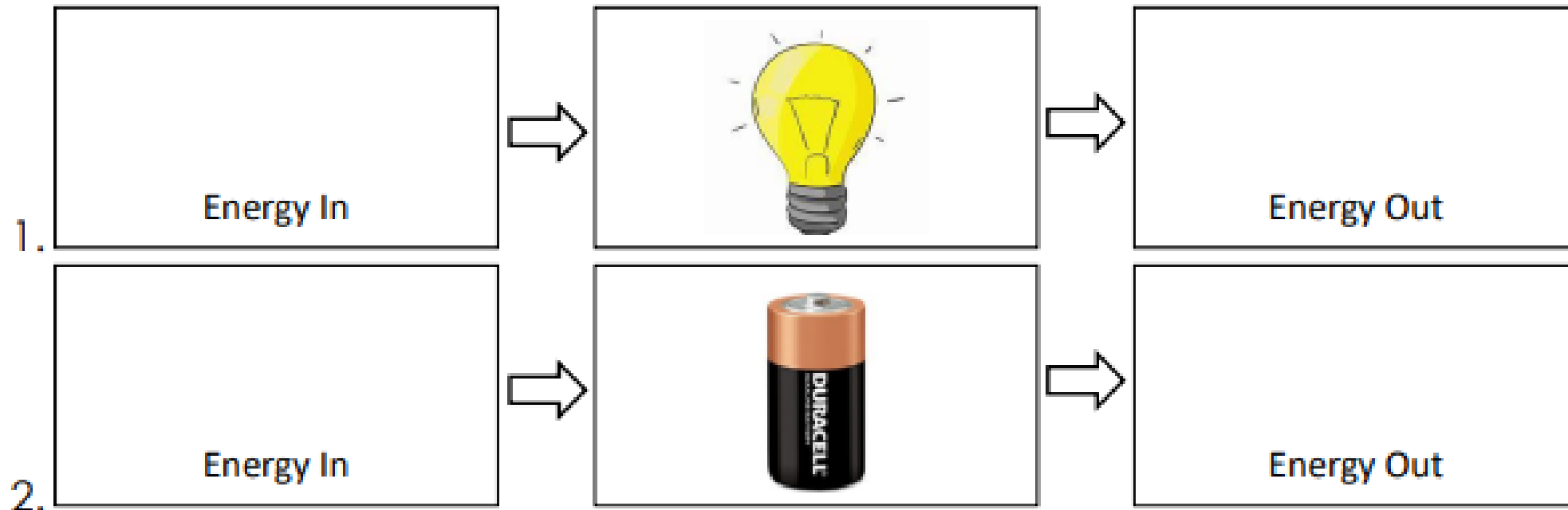
PRE-ACTIVITIES/PRE-TEST:

Energy Transformations. Identify the energies being transformed in the illustrations. Choose your answers from the box below. Some items may have multiple answers.

chemical energy
sound energy
solar energy

electrical energy
mechanical energy
light energy

heat energy
wind energy



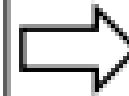
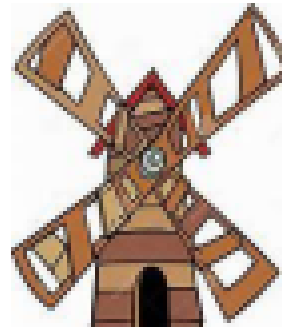
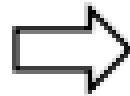
chemical energy
sound energy
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electrical energy
mechanical energy
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heat energy
wind energy

3.

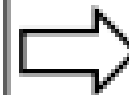
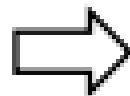
Energy In



Energy Out

4.

Energy In



Energy Out

chemical energy
sound energy
solar energy

electrical energy
mechanical energy
light energy

heat energy
wind energy

5.

Energy In



Energy Out

6.

Energy In



Energy Out

The background features abstract, overlapping geometric shapes in various shades of blue, ranging from light sky blue to deep navy blue. These shapes are primarily located on the left and right sides of the frame, creating a modern, angular aesthetic. The central area is a plain white space where the text is located.

Lesson 8: **THERMODYNAMICS**

Introduction:

- ▶ **Energy** exists in many forms, such as heat, light, chemical energy, and electrical energy. **Energy** is the ability to bring about change or to do work.
- ▶ **Thermodynamics** is the study of energy.
- ▶ Energy is fundamental entity of nature that can be transferred between parts of matter. In simplest term, **energy (E)** is the capacity to do heat (q) and work (w). Therefore, the equation is $E = q + w$

Introduction:

- ▶ It is important to study it to understand more and apply it for a certain application like driving chemical reactions.
- ▶ It is important to study energy changes in chemistry since it can be used to predict chemical reactions.

ENERGY AND ITS FORMS

▶ The **law of conservation of energy** tells us that energy is neither created nor destroyed but only transforms into other forms of energy.

Below is the list of some forms of energy:

- ▶
 - **Potential energy** is the energy by the virtue of its position. It is the stored energy.
 - **Kinetic energy** is energy in motion. Gas molecules move faster that's why they have more kinetic energy.
 - **Mechanical energy** is energy at work. A cyclist that pedals a bicycle does mechanical work.
 - **Thermal energy** is also known as heat energy. The sun emits thermal energy.
 - **Electrical energy** is the energy of the mobile electrons that produce electricity.

FIRST LAW OF THERMODYNAMICS

▶ The **Law of Conservation of Energy** states that **energy can neither be created nor destroyed. It can only be transformed from one kind to another.** The law of conservation of energy can be seen in these everyday examples of energy transfer:

Example:

▶ Water can produce electricity. Waterfalls convert potential energy to kinetic energy. This energy is then used to rotate the turbine of a generator to produce electricity. In this process, the potential energy of water in a dam can be turned into kinetic energy which can then become electric energy.



Figure 1. Maria Cristina Falls of Lanao del Norte is the primary source of electric power in Mindanao



FIRST LAW OF THERMODYNAMICS

- ▶ Example:
- ▶ Kinetic energy is changed to potential energy as a biker peddle up a hill; and potential energy is changed to kinetic energy when the bike rolls down the hill.

UNITS OF ENERGY

SI unit of energy: Joule, J

1 Joule is the amount of energy required to raise the temperature of 0.2390 g H₂O by 1°C.

$$1 \text{ kJ} = 10^3 \text{ J}$$

$$1 \text{ MJ} = 10^6 \text{ J}$$

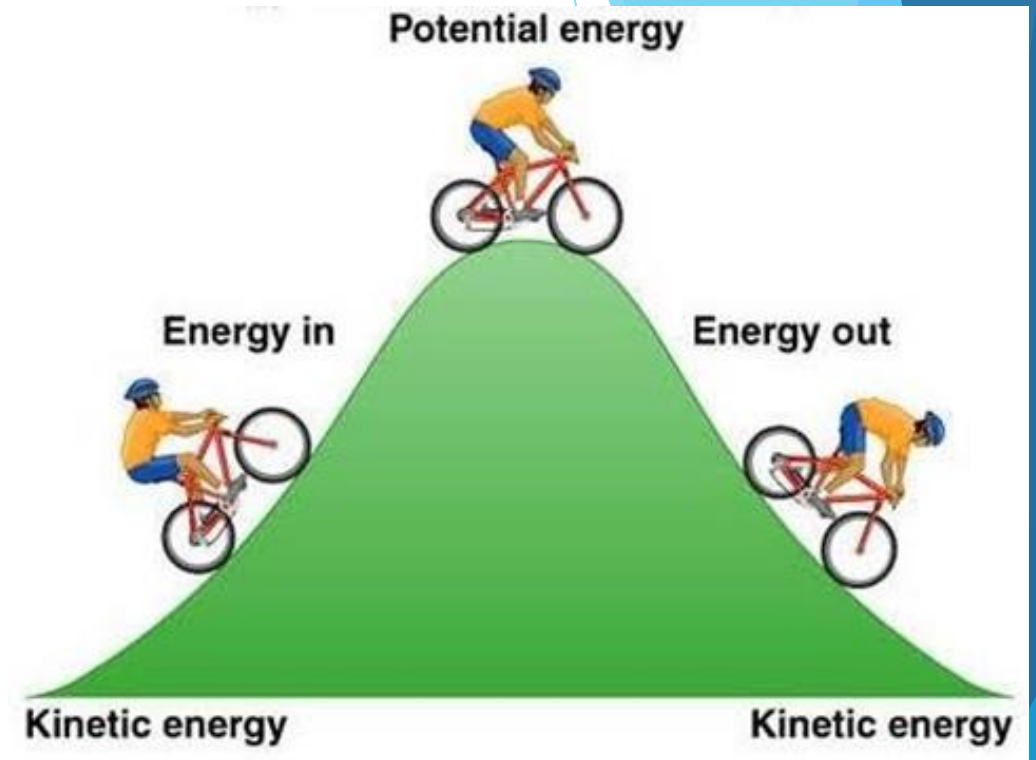
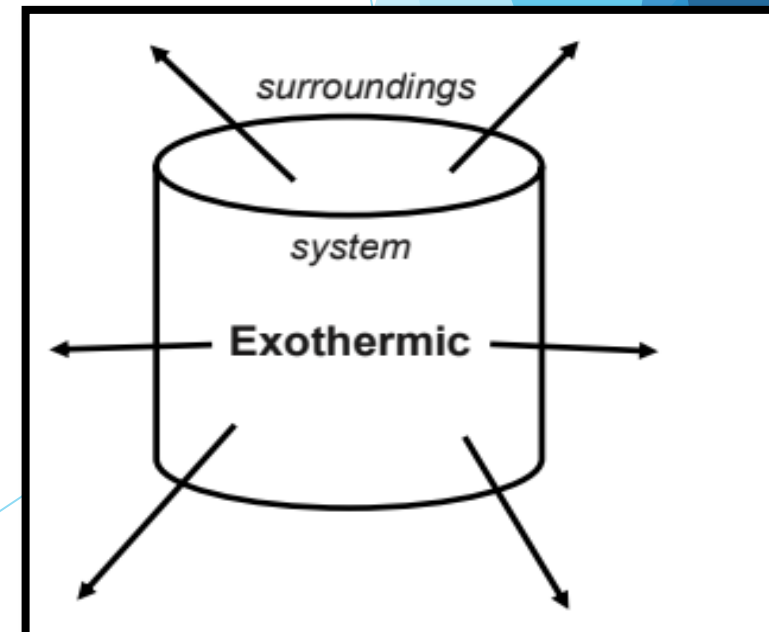
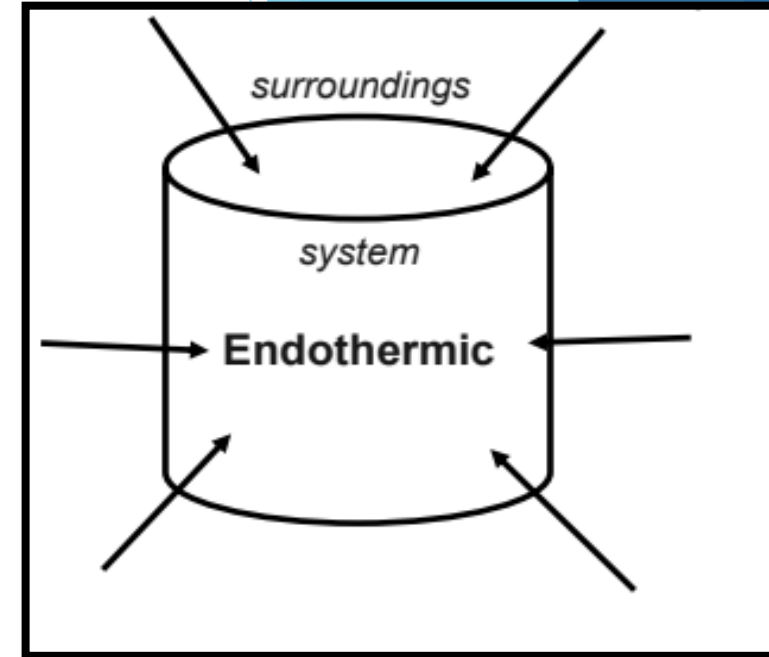


Figure 2. Interconversion of kinetic energy to potential energy.

SYSTEMS AND SURROUNDING

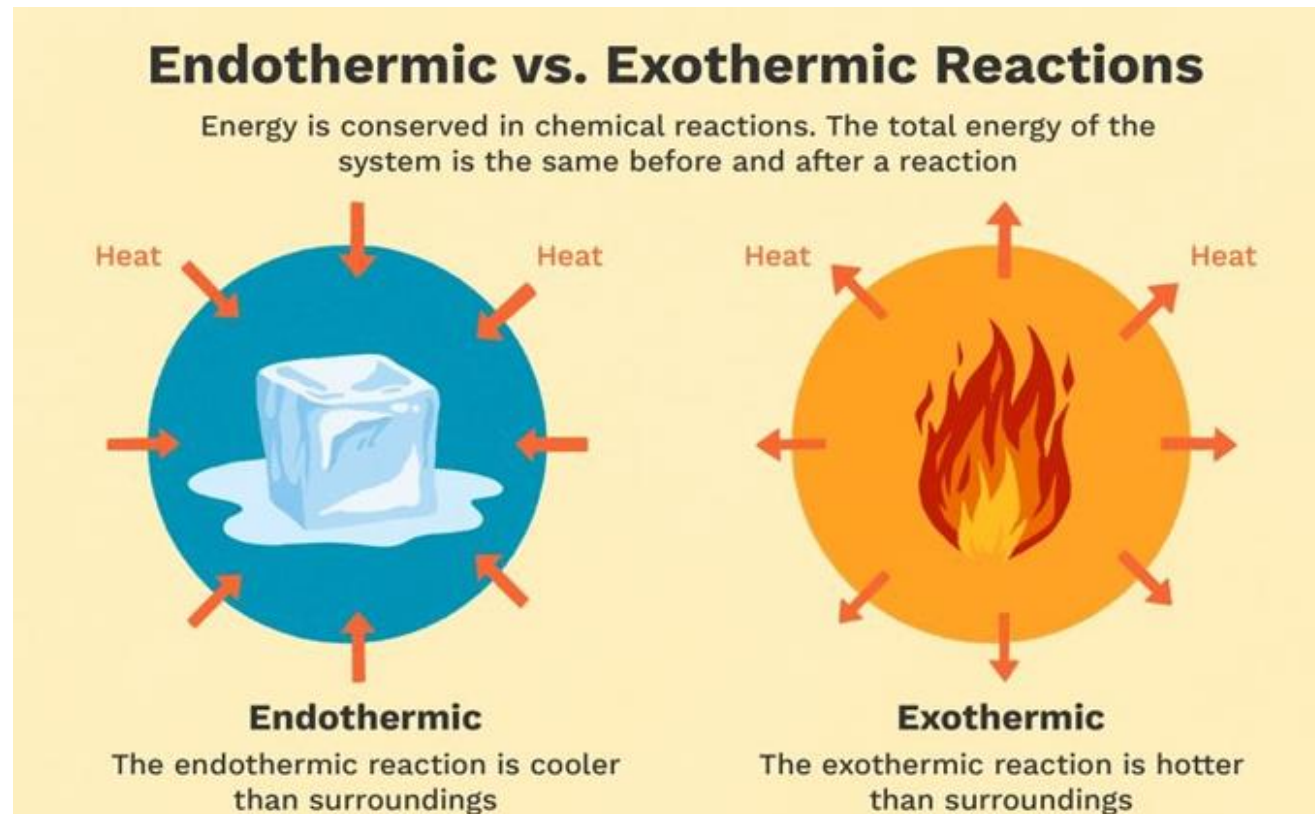
- ▶ To further understand the first law of thermodynamics, remember that the universe is often categorized into two:
- ▶ 1. the **system** which is a collection of objects that can be identified , in chemistry the system refers to a given **reaction**.
- ▶ 2. The **surrounding** which is everything else that is outside the system and it also refers to the container where the reaction takes place.



SYSTEMS AND SURROUNDING

Exothermic and Endothermic Reactions

- ▶ In every chemical reaction, there is either absorption or release of energy. Energy is required to break chemical bonds and energy is released when chemical bonds are formed.



SYSTEMS AND SURROUNDING

▶ **Endothermic Reaction**

▶ A change that involves **absorption of energy** from the surroundings is said to be an endothermic change.

▶ When an endothermic change occurs, the temperature of the reaction mixture decreases since part of the kinetic energies of particles in the surroundings are absorbed for use in the reaction.

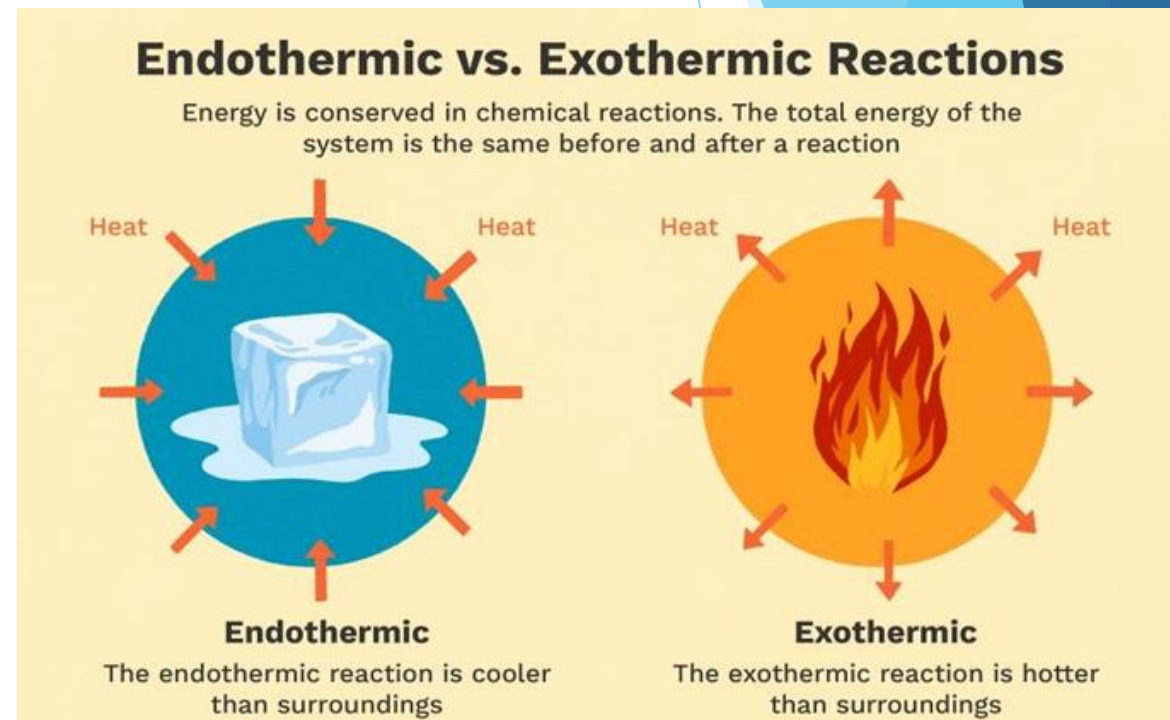


Figure 3. Endothermic VS. Exothermic

SYSTEMS AND SURROUNDING

Exothermic Reaction

If energy is released to the surroundings when a change occurs, such change is described as an exothermic change.

The **heat released** by an exothermic reaction often results to an **increase in the temperature of the reaction mixture** and the reaction vessel, and possibly the air surrounding the vessel.

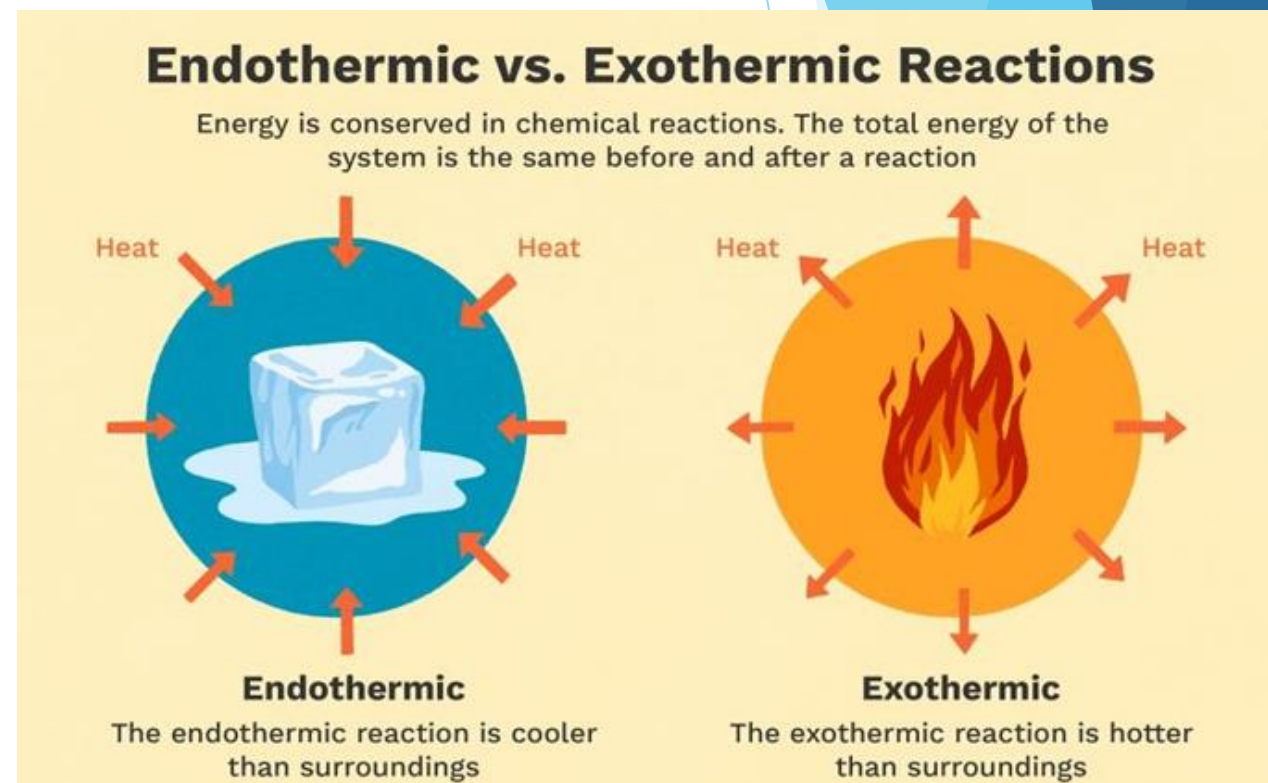


Figure 3. Endothermic VS. Exothermic

FIRST LAW OF THERMODYNAMICS

▶ In the first law of thermodynamics, the change in internal energy (ΔU) of a system is equal to the heat transfer (q) on a system and the work done (w), positive or negative, on the system.

▶ In equation:

$$\Delta U = q + w$$

▶ Here, ΔU is the change in internal energy of the system. Q is the net heat transferred into the system—that is, Q is the sum of all heat transferred into and out of the system. W is the net work done on the system.

FIRST LAW OF THERMODYNAMICS

The change in internal energy can also be determined by subtracting the final internal energy to the initial internal energy of the system.

► In equation:

► $\Delta U = U_{\text{final}} - U_{\text{initial}}$

FIRST LAW OF THERMODYNAMICS

▶ Sample Problem:

When 100,000 J of work is done on a closed system during a process, the total energy of the system increases by 55,000 J. Calculate how much heat is either added or removed from the system?

▶ Unknown: the value of q

▶ Solution:

▶ $\Delta U = q + w$ therefore, $q = \Delta U - w$

▶ $q = 55,000 \text{ J} - 100,000 \text{ J}$

▶ $= -45,000 \text{ J}$

▶ The value of q is negative which means that heat is evolved by the system or heat is removed from the system.

FIRST LAW OF THERMODYNAMICS

- ▶ Since both heat and work can be measured and quantified, it is similar in saying that change in the (E) energy of the system must result in a change in the energy of the surroundings.
- ▶ If the heat flows into a system or the surroundings do work on it (on the system), internal energy increases and the values of q and w are positive.
- ▶ On the other hand, heat flow out of the system or work done by the system (on the surroundings), the internal energy decreases because q and w are both negative.

FIRST LAW OF THERMODYNAMICS

HEAT is added to the system →	SYSTEM	→ HEAT is released by the system
(+q)		(-q)
WORK is done on the system →		→ WORK done by the system
(+w)		(-w)

SURROUNDINGS

FIRST LAW OF THERMODYNAMICS

SAMPLE PROBLEM 1:

A system absorbs 70 J of heat during a transformation. Determine the change in internal energy if the system performs 20 J of work on the surrounding.

▶ *Solution:*

▶ **the value of w is negative because the system performs work on its surroundings*

▶ $\Delta U = q + w$

▶ $= 70 \text{ J} + (-20 \text{ J})$

▶ $\Delta U = 50 \text{ J}$

▶ The amount of heat absorbed by the system during transformation is greater than the energy used to perform work. The value of ΔU is *positive* which implies that the excess heat energy caused the internal energy of the system to *increase*.

FIRST LAW OF THERMODYNAMICS

SAMPLE PROBLEM 2:

A 1500 J of heat is added to a system and 2000 J of work is done by the system. What is the change in internal energy of the system?

▶ *Solution:*

▶ **the value of w is negative because the system performs work on its surroundings*

▶ $\Delta U = q + w$

▶ $= 1500 \text{ J} + (- 2000 \text{ J})$

▶ $\Delta U = - 500 \text{ J}$

▶

The amount of heat absorbed by the system during transformation is lesser than the energy used to perform work. The value of ΔU is *negative* which implies that the internal energy of the system had *decreased*

WORK

We have seen that work can be defined as force F multiplied by distance d :

$$\blacktriangleright W = F * d$$

Where, F is the force and d is the distance.

▶ In **thermodynamics**, work can be defined in many ways such as mechanical work, electrical work and surface work. *An example of a mechanical work is the compression and expansion of a gas.* In this way, work can be defined using equation:

$$\blacktriangleright W = - P(\Delta V)$$

▶ Where, ΔV , the **change in volume**, is given by $V_f - V_i$ and P is the external atmospheric pressure.

Sample:

A certain gas expands in volume from 2.0L to 6.0L at constant temperature. Calculate the work done by the gas if it expands (a) against a vacuum and (b) against a constant pressure of 1.2 atm. The work done in gas expansion is equal to the product of the external, opposing pressure and the change in volume. What is the conversion factor between L*atm and J?

Sample:

A certain gas expands in volume from 2.0L to 6.0L at constant temperature. Calculate the work done by the gas if it expands (a) against a vacuum and (b) against a constant pressure of 1.2 atm. The work done in gas expansion is equal to the product of the external, opposing pressure and the change in volume. What is the conversion factor between L*atm and J?

(a) Because the external pressure is zero, no work is done in the expansion.

▶ Why is no work done when a gas expands against a vacuum?

▶ When a gas expands against a vacuum, no work is done because there is nothing for the gas to push against. Work is defined as the force applied over a distance, and in this case, there is no force applied because there is nothing for the gas molecules to push against.



▶ $W = -P(\Delta V)$

▶ $W = -0 (6 \text{ L} - 2 \text{ L})$

▶ $W = 0$

Sample:

A certain gas expands in volume from 2.0L to 6.0L at constant temperature. Calculate the work done by the gas if it expands (a) against a vacuum and (b) against a constant pressure of 1.2 atm. The work done in gas expansion is equal to the product of the external, opposing pressure and the change in volume. What is the conversion factor between L*atm and J?

b. The external , opposing pressure is 1.2 atm,o

- ▶ $W = -P(\Delta V)$
- ▶ $W = -1.2 \text{ atm} (6.0 \text{ L} - 2.0 \text{ L})$
- ▶ $W = -4.8 \text{ atm} \cdot \text{L}$

To convert the answers to Joules, J.

- ▶ $W = -4.8 \text{ atm} \cdot \text{L} \left(\frac{101.3 \text{ J}}{1 \text{ atm} \cdot \text{L}} \right)$
- ▶ $W = -4.9 \times 10^2 \text{ J}$

Because this is gas expansion (work is done by the system on the surroundings), the work done has a negative sign.

HEAT

- ▶ The other component of internal energy is **heat, Q** . For example, it takes 4184 J of energy to raise the temperature of 100g of water from 20°C to 30°C. This energy can be gained directly as
 - ▶ (a) heat energy from a Bunsen burner, without doing any work on the water;
 - ▶ (b) by doing work on the water without adding heat energy (for example, by stirring the water with a magnetic stir bar); or
 - ▶ (c) by some combination of the procedures described in (a) and (b).
- ▶ **This simple illustration shows that heat associated with a given process, like work, depends on how the process is carried out.**
- ▶ It is important to note that regardless of which procedure is taken, the change in internal energy of the system, ΔU , depends on the sum of $(Q + W)$.

SAMPLE PROBLEM:

- ▶ If a system has 50J of heat added to it, and it does 20J of work, what is the change in its internal energy?

SAMPLE PROBLEM:

If a system has 50J of heat added to it, and it does 20J of work, what is the change in its internal energy?

$$\text{Given: } q = 50\text{J}$$

$$w = -20\text{J}$$

$$\text{Unknown: } \Delta U$$

$$\text{Formula: } \Delta U = q + w$$

$$\text{Solution: } \Delta U = 50\text{J} + (-20\text{J})$$

$$\boxed{\Delta U = 30\text{J}}$$

SAMPLE PROBLEM:

- ▶ (a) A gas at expands from 2.0 L to 6.0 L at a constant pressure of 912 mmHg. If q was zero in the process, what would be the change in internal energy?

SAMPLE PROBLEM:

(a) A gas at expands from 2.0 L to 6.0 L at a constant pressure of 912 mmHg. If q_p was zero in the process, what would be the change in internal energy?

Given: $V_F = 6.0 \text{ L}$

$V_I = 2.0 \text{ L}$

$P = 912 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 1.2 \text{ atm}$

$q = 0$

Unknown: ΔU

Formula: $\Delta U = q + w$
 $= q + (-P\Delta V)$

Solution: $\Delta U = 0 \text{ J} - 1.2 \text{ atm} (6.0 \text{ L} - 2.0 \text{ L})$

$= 0 \text{ J} - 1.2 \text{ atm} (4.0 \text{ L})$

$= 0 \text{ J} - 4.8 \text{ atm} \cdot \text{L}$

$= -4.8 \text{ atm} \cdot \text{L} \times \frac{101.3 \text{ J}}{1 \text{ atm} \cdot \text{L}} =$

$\boxed{-486.24 \text{ J}}$

ENTHALPY

- ▶ The term **enthalpy** originates from the Greek word **enthalpein** that means “**to warm**”.
- ▶ Enthalpy is a thermodynamic property of a system. It is the sum of the **internal energy added to the product of the pressure and volume of the system.**

ENTHALPY

- ▶ Enthalpy is denoted as **H**; specific enthalpy denoted as **h**. Enthalpy in a throttling process is constant. We can define enthalpy using equation.

- ▶
$$H = U + PV$$

- ▶ Where, **H** is the enthalpy of a reaction, **U** is the internal energy of the system, **P** is the pressure and **V** is the volume of the system.
- ▶ Because U and PV have energy units, enthalpy also has energy units. Common units used to express enthalpy are the **Joule, calorie, or BTU (British Thermal Unit.)**

ENTHALPY

- ▶ **Change in enthalpy** is calculated rather than enthalpy in part because total enthalpy of a system cannot be measured since it is impossible to know the zero point.
- ▶ Because most reactions are constant-pressure processes, we can equate the heat change in these cases, to the change in enthalpy. For any reaction of the type
 - ▶ **reactants \rightarrow products**

ENTHALPY

- ▶ We define the change in enthalpy, called the **enthalpy of reaction**, ΔH_{rxn} , as the difference between the enthalpies of the products and the enthalpies of the reactants. It represents heat absorbed or released when the reactants are converted into products, at a constant pressure.

- ▶
$$\Delta H_{\text{rxn}} = H_{\text{(Products)}} - H_{\text{(Reactants)}}$$

- ▶ Remember that the enthalpy of reaction can be positive or negative, depending on the process. For an **endothermic process** (heat absorbed by the system from the surroundings), ΔH is **positive (that is, $\Delta H > 0$)**. For an **exothermic process** (heat released by the system to the surroundings), ΔH is **negative (that is, $\Delta H < 0$)**.

ENTHALPY

- ▶ When the system changes at **constant pressure**, the change in enthalpy, ΔH , is
 - ▶ $\Delta H = \Delta(U + PV)$
- ▶ This can be written
 - ▶ $\Delta H = \Delta U + P\Delta V$
- ▶ **Since** $\Delta U = q + w$ **and** $w = -P\Delta V$, we can substitute these into the enthalpy expression:
 - ▶ $\Delta H = \Delta U + P\Delta V$
 - ▶ $\Delta H = (q+w) - w$
 - ▶ $\Delta H = q$
- ▶ The **enthalpy change**, ΔH , is defined as the heat gained or lost by the system under constant pressure.
 - ▶ $\Delta H = q_p$

STANDARD ENTHALPHY OF VARIOUS SUBSTANCES

Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mol)
AgBr(s)	-99.5	C ₂ H ₂ (g)	+226.7
AgCl(s)	-127.0	C ₂ H ₄ (g)	+52.3
AgI(s)	-62.4	C ₂ H ₆ (g)	-84.7
Ag ₂ O(s)	-30.6	C ₃ H ₈ (g)	-103.8
Ag ₂ S(s)	-31.8	n-C ₄ H ₁₀ (g)	-124.7
Al ₂ O ₃ (s)	-1669.8	n-C ₅ H ₁₂ (l)	-173.1
BaCl ₂ (s)	-860.1	C ₂ H ₅ OH(l)	-277.6
BaCO ₃ (s)	-1218.8	CoO(s)	-239.3
BaO(s)	-558.1	Cr ₂ O ₃ (s)	-1128.4
BaSO ₄ (s)	-1465.2	CuO(s)	-155.2
CaCl ₂ (s)	-795.0	Cu ₂ O(s)	-166.7
CaCO ₃	-1207.0	CuS(s)	-48.5

CaO(s)	-635.5	CuSO ₄ (s)	-769.9
Ca(OH) ₂ (s)	-986.6	Fe ₂ O ₃ (s)	-822.2
CaSO ₄ (s)	-1432.7	Fe ₃ O ₄ (s)	-1120.9
CCl ₄ (l)	-139.5	HBr(g)	-36.2
CH ₄ (g)	-74.8	HCl(g)	-92.3
CHCl ₃ (l)	-131.8	HF(g)	-268.6
CH ₃ OH(l)	-238.6	HI(g)	+25.9
CO(g)	-110.5	HNO ₃ (l)	-173.2
CO ₂ (g)	-393.5	H ₂ O(g)	-241.8
H ₂ O(l)	-285.8	NH ₄ Cl(s)	-315.4
H ₂ O ₂ (l)	-187.6	NH ₄ NO ₃ (s)	-365.1
H ₂ S(g)	-20.1	NO(g)	+90.4
H ₂ SO ₄ (l)	-811.3	NO ₂ (g)	+33.9

STANDARD ENTHALPHY OF VARIOUS SUBSTANCES

HgO(s)	-90.7	NiO(s)	-244.3
HgS(s)	-58.2	PbBr ₂ (s)	-277.0
KBr(s)	-392.2	PbCl ₂ (s)	-359.2
KCl(s)	-435.9	PbO(s)	-217.9
KClO ₃ (s)	-391.4	PbO ₂ (s)	-276.6
KF(s)	-562.6	Pb ₃ O ₄ (s)	-734.7
MgCl ₂ (s)	-641.8	PCl ₃ (g)	-306.4
MgCO ₃ (s)	-1113	PCl ₅ (g)	-398.9
MgO(s)	-601.8	SiO ₂ (s)	-859.4
Mg(OH) ₂ (s)	-924.7	SnCl ₂ (s)	-349.8
MgSO ₄ (s)	-1278.2	SnCl ₄ (l)	-545.2
MnO(s)	-384.9	SnO(s)	-286.2

MnO ₂ (s)	-519.7	SnO ₂ (s)	-580.7
NaCl(s)	-411.0	SO ₂ (g)	-296.1
NaF(s)	-569.0	So ₃ (g)	-395.2
NaOH(s)	-426.7	ZnO(s)	-348.0
NH ₃ (g)	-46.2	ZnS(s)	-202.9

<https://www2.chem.wisc.edu/deptfiles/genchem/netorial/modules/thermodynamics/table.htm>

ENTHALPY of REACTION

▶ Sample Problem:



▶ Given: ΔH_f° for $\text{H}_2\text{O} = -241.8 \text{ kJ/mol}$

▶ ΔH_f° for $\text{H}_2 = 0$

▶ ΔH_f° for $\text{O}_2 = 0$

$$\Delta H_{\text{rxn}} = \Sigma H_{(\text{Products})} - \Sigma H_{(\text{Reactants})}$$

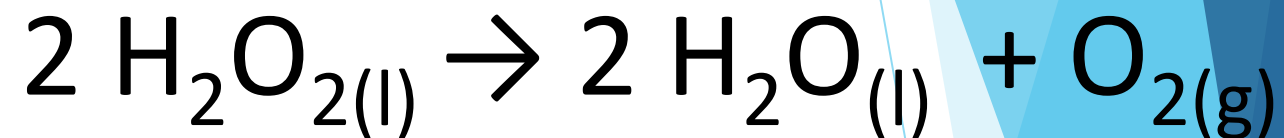
▶ $\Delta H_{\text{rxn}} = [2(0) + (0)] - [2(-241.8)]$

▶ $\Delta H_{\text{rxn}} = [0] - [-483.6]$

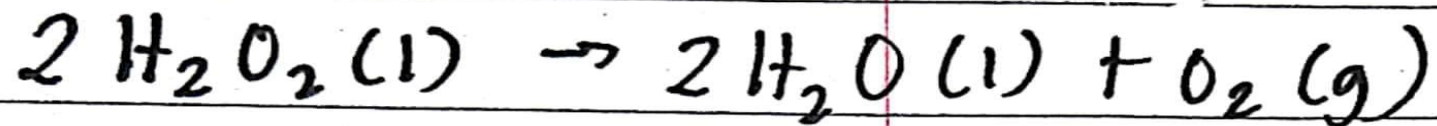
▶ $\Delta H_{\text{rxn}} = 483.6 \text{ kJ} - \text{Endothermic}$

CaO(s)	-635.5	CuSO ₄ (s)	-769.9
Ca(OH) ₂ (s)	-986.6	Fe ₂ O ₃ (s)	-822.2
CaSO ₄ (s)	-1432.7	Fe ₃ O ₄ (s)	-1120.9
CCl ₄ (l)	-139.5	HBr(g)	-36.2
CH ₄ (g)	-74.8	HCl(g)	-92.3
CHCl ₃ (l)	-131.8	HF(g)	-268.6
CH ₃ OH(l)	-238.6	HI(g)	+25.9
CO(g)	-110.5	HNO ₃ (l)	-173.2
CO ₂ (g)	-393.5	H ₂ O(g)	-241.8
H ₂ O(l)	-285.8	NH ₄ Cl(s)	-315.4
H ₂ O ₂ (l)	-187.6	NH ₄ NO ₃ (s)	-365.1
H ₂ S(g)	-20.1	NO(g)	+90.4
H ₂ SO ₄ (l)	-811.3	NO ₂ (g)	+33.9

► Calculate the change in enthalpy of the reaction.



Calculate the change in enthalpy of the reaction:



$$2(-187.6) \rightarrow 2(-285.8) + 0$$

$$\Delta H_{\text{rxn}} = \sum H_{\text{products}} - \sum H_{\text{reactants}}$$

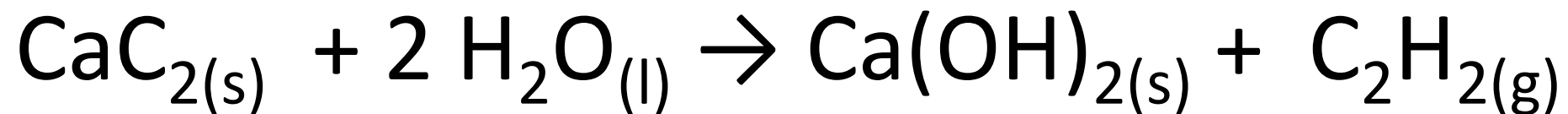
$$= [2 \text{ mol}(-285.8 \text{ kJ/mol}) + 0] - [2 \text{ mol}(-187.6 \text{ kJ/mol})]$$

$$= [-571.6] - [-375.2]$$

$$= -571.6 \text{ kJ} + 375.2 \text{ kJ}$$

$$\boxed{\Delta H_{\text{rxn}} = -196.4 \text{ kJ}}$$

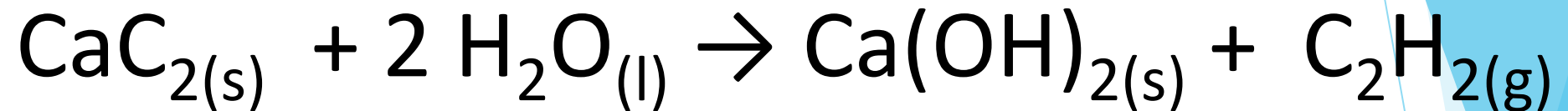
► Calculate the change in enthalpy of the reaction.



The Standard Enthalpy and Entropy of Various Substances

Substance	ΔH_f° (kJ/mol)	S° (J/K · mol)
$\text{C}_4\text{H}_{10(g)}$	-126	310
$\text{CaC}_{2(s)}$	-63	70.
$\text{Ca(OH)}_{2(s)}$	-987	83
$\text{C}_2\text{H}_{2(g)}$	227	201
$\text{CO}_{2(g)}$	-394	214
$\text{H}_{2(g)}$	0	131
$\text{H}_2\text{O}_{(g)}$	-242	189
$\text{H}_2\text{O}_{(L)}$	-286	70.
$\text{NH}_{3(g)}$	-46	193
$\text{NO}_{(g)}$	90.	211
$\text{NO}_{2(g)}$	34	240.
$\text{N}_2\text{O}_{(g)}$	82	220.
$\text{O}_{2(g)}$	0	205
$\text{O}_{3(g)}$	143	239

- Calculate the change in enthalpy of the reaction.



$$\Delta H_{\text{rxn}} = \sum H_{(\text{Products})} - \sum H_{(\text{Reactants})}$$

$$\Delta H_{\text{rxn}} = [-987 + 227] - [-63 + 2(-286)]$$

$$\Delta H_{\text{rxn}} = [-760] + [-635]$$

$$\Delta H_{\text{rxn}} = -125 \text{ kJ}$$

The Standard Enthalpy and Entropy of Various Substances

Substance	ΔH_f° (kJ/mol)	S° (J/K · mol)
$\text{C}_4\text{H}_{10(g)}$	-126	310
$\text{CaC}_{2(s)}$	-63	70.
$\text{Ca(OH)}_{2(s)}$	-987	83
$\text{C}_2\text{H}_{2(g)}$	227	201
$\text{CO}_{2(g)}$	-394	214
$\text{H}_{2(g)}$	0	131
$\text{H}_2\text{O}_{(g)}$	-242	189
$\text{H}_2\text{O}_{(L)}$	-286	70.
$\text{NH}_{3(g)}$	-46	193
$\text{NO}_{(g)}$	90.	211
$\text{NO}_{2(g)}$	34	240.
$\text{N}_2\text{O}_{(g)}$	82	220.
$\text{O}_{2(g)}$	0	205
$\text{O}_{3(g)}$	143	239

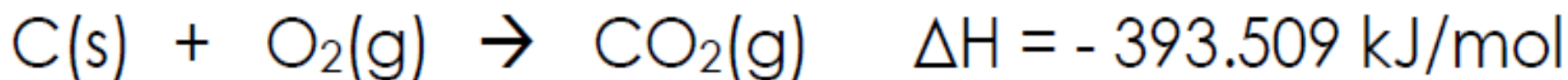
HESS LAW

- ▶ Another way of calculating the change of enthalpy is by using **Hess Law**. Hess Law states that the enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs.
- ▶ In other words, Hess Law is saying that if you convert reactants A into products B, the **overall enthalpy change** will be the **same** whether you do it in one step or two steps or however many steps.
- ▶ According to Hess's Law that **regardless the multiple stages or steps of a reaction, the total enthalpy change for the reaction is sum of all changes**.

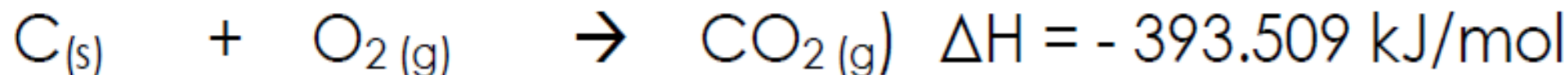
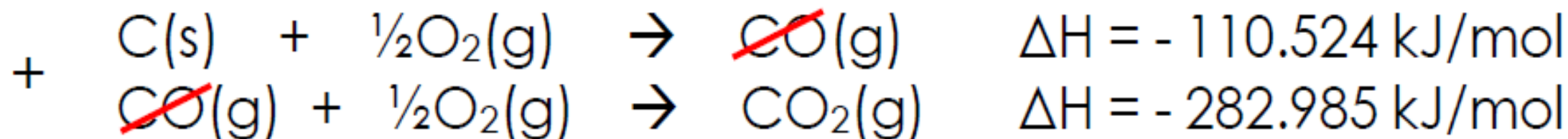
Example:

For example,

Carbon can be burned directly to carbon dioxide.



Or it can be done in two steps: carbon to carbon monoxide then carbon monoxide to carbon dioxide.



Example:

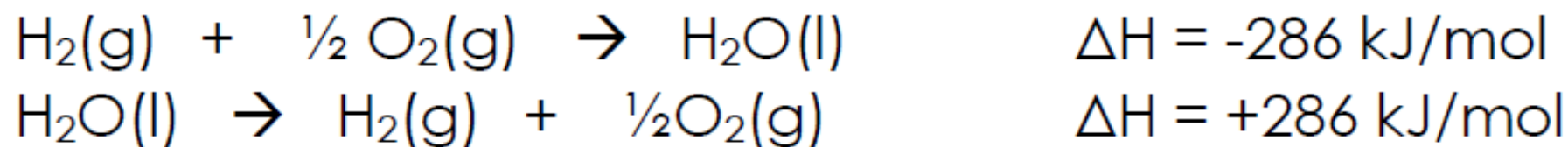
- Cancel CO because it appears on both sides of the equations.
- From this example, the overall change is the net result of a series of steps, and the net value of ΔH for the overall reaction is just the sum of all the enthalpy changes of the different steps.
- Note that in the above reaction the heats of reaction of the individual steps involved are added algebraically to obtain the overall heat of reaction.

► Thermochemical equations possess two properties:

1. They may be reversed.
2. They may be treated as algebraic expressions. Therefore, they may be added, subtracted, multiplied by a factor or divided by a factor.

1. Reversing thermochemical equations

When we reverse an equation the sign of ΔH is also reversed.



This means that the heat involved in the formation of one mole of $\text{H}_2\text{O}(\text{l})$ is equal to the amount of heat required to decompose one mole of liquid water.

► **Thermochemical equations possess two properties:**

2. They may be treated as algebraic expressions.

Therefore, they may be added, subtracted, multiplied by a factor or divided by a factor.

The equation involved is



► **Thermochemical equations possess two properties:**

2.They may be treated as algebraic expressions. Therefore, they may be added, subtracted, multiplied by a factor or divided by a factor.

To solve the problem, combine equations (a), (b) and (c) so that when added, everything cancels out except the formulas in the desired equation, that is the formation of methane.

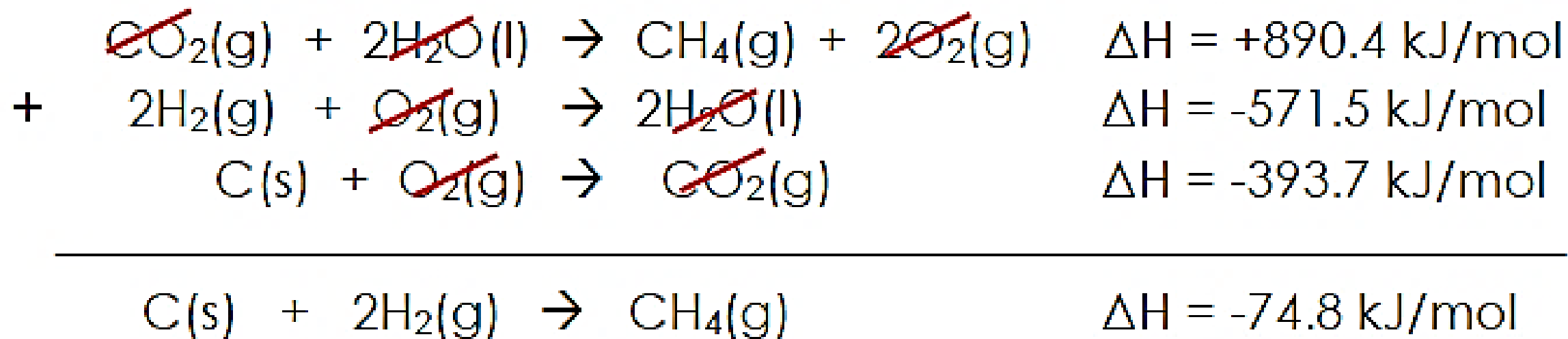
Note that in the desired chemical equation CH₄ is on the product side, thus reverse equation (a) and also reverse the sign of ΔH.



► **Thermochemical equations possess two properties:**

2.They may be treated as algebraic expressions. Therefore, they may be added, subtracted, multiplied by a factor or divided by a factor.

There is no need to do anything with equations (b) and (c) since $C_{(s)}$ and $H_{2(g)}$ are on the reactant side. In the desired equation, they are also on the reactants side. So add the three equations, cancelling terms that appear on both sides:



► **To summarize how to calculate for the ΔH of a specific reaction:**

1. Look for the formulas that **appear only once among the equations** and place it in the right place just as that in the desired equation.

2. Note the number of moles of each reactant and products in the desired equation.

▶ To summarize how to calculate for the ΔH of a specific reaction:

3. Manipulate the equations with known ΔH values so that the number of moles of reactants and products are on the correct sides as in the desired equation. Do not forget to

▶ Change the sign of ΔH when the equation is reversed.

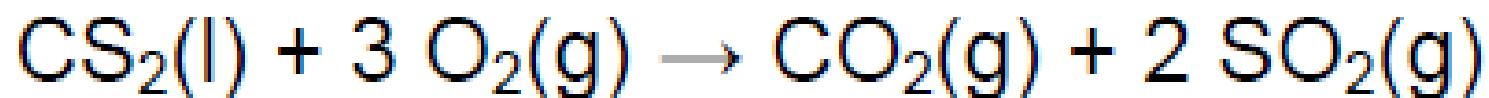
▶ Multiply/divide the number of moles and ΔH by the same factor.

► To summarize how to calculate for the ΔH of a specific reaction:

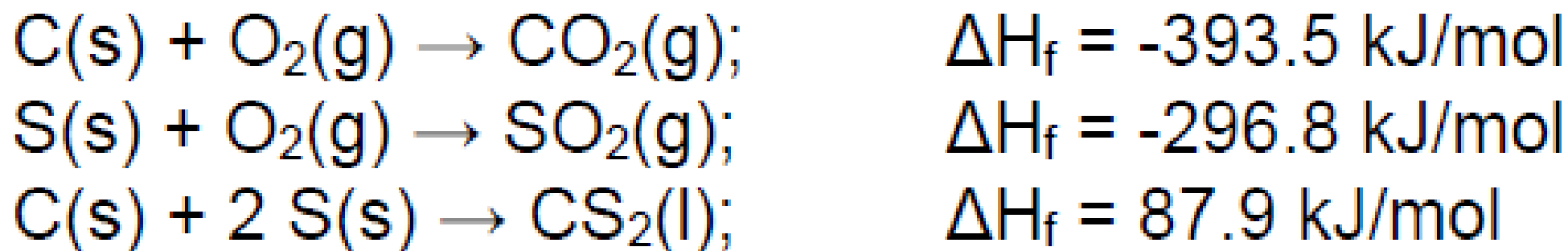
4. **Add** the manipulated equations,
cancelling terms that are common to both sides of the equation to obtain the desired equation.

Algebraically add the ΔH values to get the final ΔH or change in enthalpy of the desired equation.

Example 3.4.1. Find the enthalpy change for the reaction.



when:



Example 3.4.1. Find the enthalpy change for the reaction.



when:

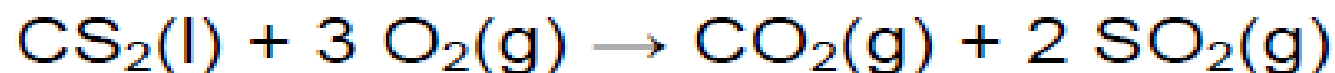


This is an example of a trial-and-error problem. To solve this, it is better to start with a reaction with only one mole of reactant or product in the reaction.

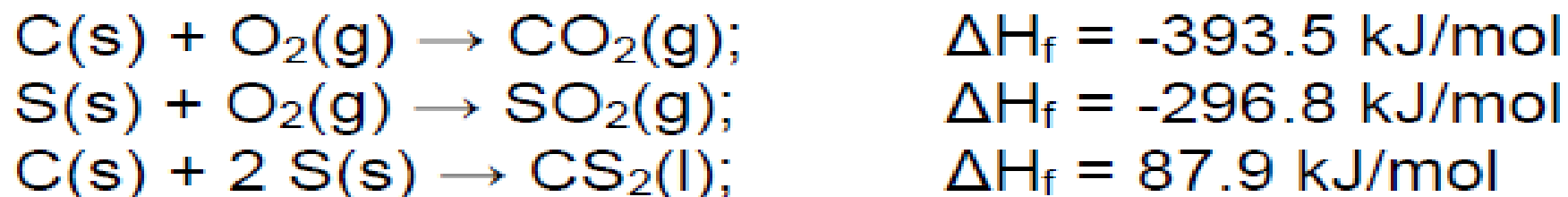
Our reaction needs one CO_2 in the product and the first reaction also has one CO_2 product.



Example 3.4.1. Find the enthalpy change for the reaction.



when:



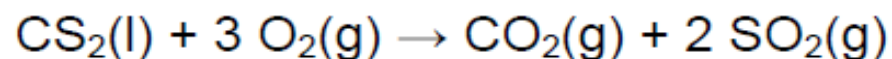
This reaction gives us the CO_2 needed on the product side and one of the O_2 needed on the reactant side. The other two O_2 can be found in the second reaction.



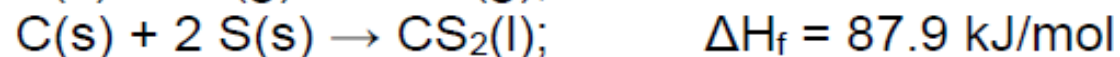
Since only one O_2 is in the reaction, multiply the reaction by two to get the second O_2 . This doubles the ΔH_f value.



Example 3.4.1. Find the enthalpy change for the reaction.



when:



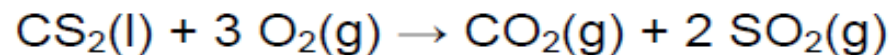
Combining these equations gives



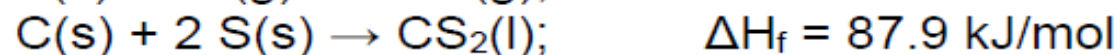
The enthalpy change is the sum of the two reactions:

$$\Delta H_f = -393.5 \text{ kJ/mol} + -593.6 \text{ kJ/mol} = -987.1 \text{ kJ/mol}$$

Example 3.4.1. Find the enthalpy change for the reaction.



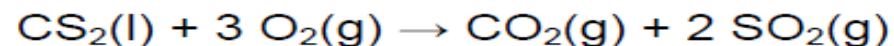
when:



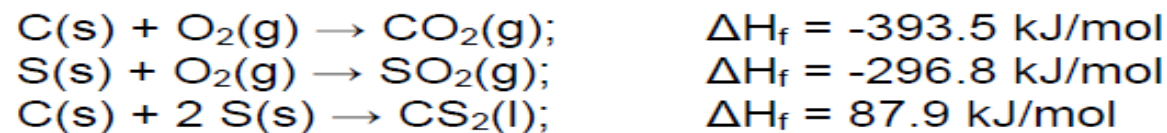
This equation has the product side needed in the problem but contains an extra two S and one C atom on the reactant side. Fortunately, the third equation has the same atoms. If the reaction is reversed, these atoms are on the product side. When the reaction is reversed, the sign of the change in enthalpy is reversed.



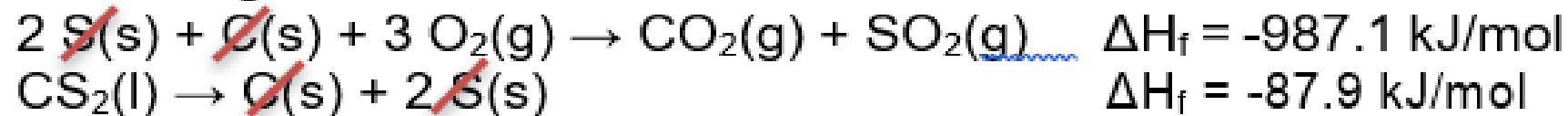
Example 3.4.1. Find the enthalpy change for the reaction.



when:



Add these two reactions together and the extra S and C atoms cancel out. The remaining reaction is the reaction needed in the question. Since the reactions were added together, their ΔH_f values are added together.

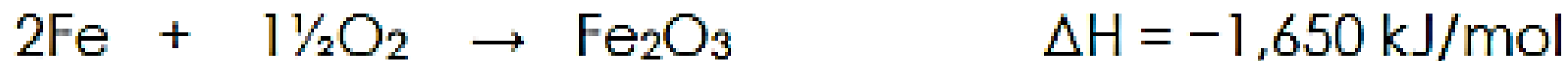


Therefore, the change in enthalpy for the reaction is -1075 kJ/mol.

SAMPLE PROBLEM: HESS LAW



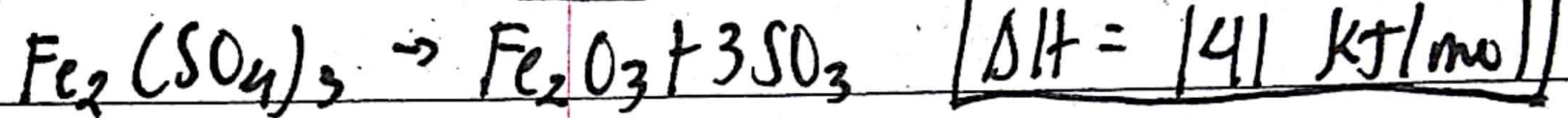
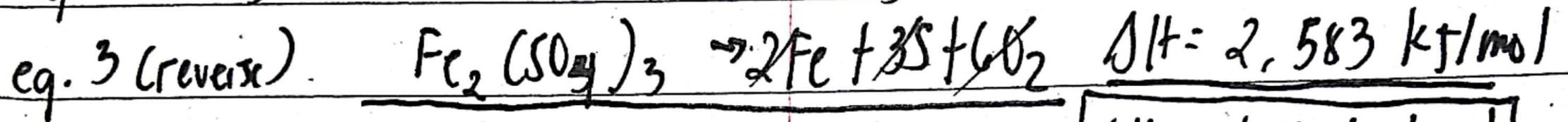
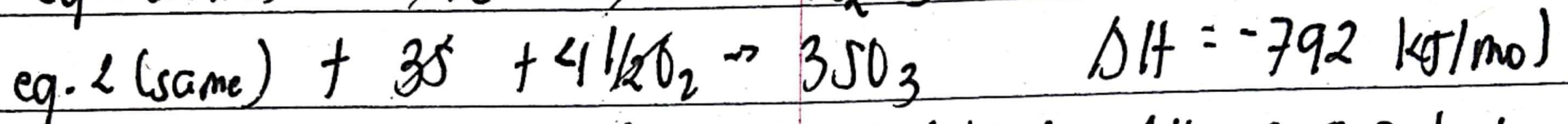
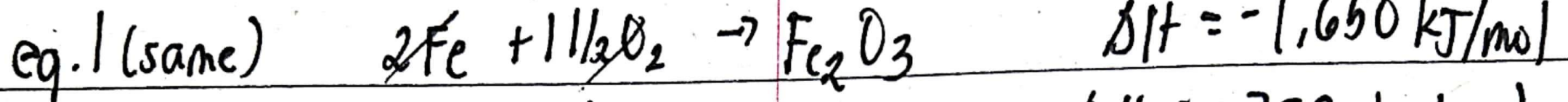
given these data:



SAMPLE PROBLEM: HESS LAW



given these data:



SAMPLE PROBLEM: HESS LAW



given these data:



SAMPLE PROBLEM: HESS LAW



$$\Delta H = \underline{\hspace{2cm}}$$

given these data:

