Name types of Energy:

All the “forms” of energy … Light, Nuclear, Mechanical, Heat, Chemical, Electrical, Magnetic, and Sound involve the interaction of \_\_\_\_\_\_\_ & \_\_\_\_\_\_\_.

Energy

* The capacity to do work (“\_\_\_\_\_\_\_” 🡪 KJ)
* Unlike matter, energy has neither mass nor volume. Energy is detected only by its effects.

\_\_\_\_\_\_\_ Energy (PE) 🡪 \_\_\_\_\_\_\_ energy

* Every chemical substance contains a certain amount of energy, called chemical potential energy, which is stored in the chemical \_\_\_\_\_\_\_ of its \_\_\_\_\_\_\_.
* The kinds of atoms and their arrangement in a substance, determine the amount of potential energy stored in that substance.

\_\_\_\_\_\_\_ Energy (KE) 🡪 energy of \_\_\_\_\_\_\_

* \_\_\_\_\_\_\_ is a measure of the \_\_\_\_\_\_\_ kinetic energy of the particles in a sample of matter. The greater the \_\_\_\_\_\_\_ of the particles in a sample, the hotter it feels.
* Celsius and Kelvin temperatures are related by the following equation:
* \_\_\_ = \_\_\_\_\_\_\_ + ⁰\_\_\_

Energy *associated with chemical changes and with physical changes*

* \_\_\_\_\_\_\_ ΔH = \_\_\_
	+ heat flows \_\_\_\_\_\_\_ the system from the surroundings
	+ Exo\_\_\_\_\_\_\_ thermic ΔH = \_\_\_
		- heat flows \_\_\_\_\_\_\_ the system into the surroundings
	+ \_\_\_\_\_\_\_ energy 🡪 The energy \_\_\_\_\_\_\_ before a chemical reaction can take place (e.g. Powder needs to be airbourne before it can explode.)

Energy Units

* \_\_\_\_\_\_\_ (J) 🡪 The joule (J) is the SI unit of energy. One joule of heat raises the temperature of 1 g of pure water 0.2390 ⁰C.
* \_\_\_\_\_\_\_ (C) 🡪 *The amount of energy required to raise 1 gram of water, one degree Celsius*.
* *1 calorie = 4.18 joules [Reference Table 1]*
* *1 J = 2.39 X 10-1 cal*

Heat

* A form of energy that takes into account the quantity of \_\_\_\_\_\_\_
Units: calories (Kcal) or joules (KJ)
* Calorimeter 🡪 instrument used to measure heat
* Physical and chemical changes involve changes in energy:
* Exothermic - \_\_\_\_\_\_\_ energy or heat
* Endothermic - \_\_\_\_\_\_\_ /\_\_\_\_\_\_\_ energy or heat
* \_\_\_\_\_\_\_ The measure of heat intensity: describes the average \_\_\_\_\_\_\_ energy (KE) of the molecules in a system
* Units: measured in degrees
	+ Fahrenheit (°F)
	+ Celsius (°C); also called centigrade
* \_\_\_\_\_\_\_ (K)
* “\_\_\_\_\_\_\_ temperature”
* Begins with \_\_\_ K (*theoretical temperature*)
* includes \_\_\_\_\_\_\_ and \_\_\_\_\_\_\_
* Conversions
	+ Fahrenheit and Celsius

F = 9/5 C + 32

C = 5/9 (F - 32)

* Kelvin vs. Celsius

\_\_\_ = \_\_\_ + \_\_\_\_\_\_\_

1 K increment = 1 C

Temperature versus Heat

* \_\_\_\_\_\_\_ involves the motion of molecules only
* Heat incorporates temperature and \_\_\_\_\_\_\_ (amount)
* Example: *burning a match verses a bonfire*
	+ *Temperature for both is the same (~400-600 °C)*
	+ *The bonfire contains MUCH more heat*

Heat flows from \_\_\_\_\_\_\_ to \_\_\_\_\_\_\_ 🡪 This represents K\_\_\_\_\_\_\_ E

Heating and Cooling Curves

Why doesn’t the temperature change as ice melts? As water boils?

What kind of energy relationships are occurring as ice melts and then boils?

What do we call the stages when ice melts and water boils?

Phase Change Diagrams

\_\_\_\_\_\_\_ ΔH = + \_\_\_\_\_\_\_ Curve

Heat flows \_\_\_\_\_\_\_ the system from the surroundings

Label KE/PE, phase, heat equation



|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Graph | Temp | Phase | KE | PE |
|  | A 🡪 B | -20 🡪 0 | s |  |  |
|  | B 🡪 C | 0 | s 🡪 l |  |  |
|  | C 🡪 D | 0 🡪 100 | l |  |  |
|  | D 🡪 E | 100 | l 🡪 g |  |  |
|  | E 🡪 F | 100 🡪 110 | g |  |  |

\_\_\_\_\_\_\_ ΔH = ─ \_\_\_\_\_\_\_ Curve

Heat flows \_\_\_\_\_\_\_ the system into the surroundings

Label KE/PE, phase, Heat equation

Heat

Temperature

°C

?

?

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Temp | Phase | KE | PE |
|  | 120 🡪 100 | s |  |  |
|  | 100 | s 🡪 l |  |  |
|  | 100 🡪 0 | l |  |  |
|  | 0 | l 🡪 g |  |  |
|  | 0 🡪 -10 | g |  |  |

General Heating Curve

Fixed points (based on PE) … \_\_\_\_\_\_\_ is “fixed”

* \_\_\_\_\_\_\_ point ↔ \_\_\_\_\_\_\_ point (s ↔ l)
* \_\_\_\_\_\_\_ point ↔ \_\_\_\_\_\_\_ pt (l ↔ g)
* Sublimation ↔ Deposition (s ↔ g)

Quantitative Heat Measurements

Heat Equation: q = m c ∆T \_\_\_\_\_\_\_ (type of energy)

q: change in \_\_\_\_\_\_\_ energy for system or surroundings

m: \_\_\_\_\_\_\_ of substance

c: \_\_\_\_\_\_\_ heat of substance

cp (H20): 1.0 cal/g∙⁰C = 4.18 j/g∙⁰C

c (ice): 0.5cal/g∙⁰C = 2.09 j/g∙⁰C

c (steam): 0.5cal/g∙⁰C = 2.09 j/g∙⁰C

∆ T: change in \_\_\_\_\_\_\_ (closed system)

∆ T = \_\_\_\_\_\_\_ Temp – \_\_\_\_\_\_\_ Temp … ∆ T = Tf – Ti

Heat of \_\_\_\_\_\_\_ / Solidification (\_\_\_\_\_\_\_)

* The amount of energy gained or released when a substance \_\_\_\_\_\_\_ (s 🡪 l) or \_\_\_\_\_\_\_ (l 🡪 s)
* Melting (s 🡪 l): \_\_\_\_\_\_\_ ∆Hf = +
* Freezing (l 🡪 s): \_\_\_\_\_\_\_ ∆Hf = -

Heat of \_\_\_\_\_\_\_ / Condensation (PE)

* The amount of energy gained or released when a substance \_\_\_\_\_\_\_ (l 🡪 g) or \_\_\_\_\_\_\_ (g 🡪 l)
* vaporizing (l 🡪 g): \_\_\_\_\_\_\_ ∆Hv = +
* condensing (g 🡪 l): \_\_\_\_\_\_\_ ∆Hv = -

Specific Heat of Water

* The \_\_\_\_\_\_\_ heat capacity, or simply the specific heat (C) of a substance, is the amount of heat it takes to raise the temperature of 1 g of the substance 1⁰ C.
* \_\_\_\_\_\_\_ has the second highest specific heat of all liquids.
* Metals generally have \_\_\_\_\_\_\_ specific heats.
* Some things heat up or cool down faster than others.
	+ \_\_\_\_\_\_\_ heats up and cools down faster than \_\_\_\_\_\_\_
	+ e.g. Michigan stays warmer in the winter due to the heat from the lakes (*e.g. hot vegetables stay hot a long time because of their high water content*)
* Specific heat is ability of a substance to “\_\_\_\_\_\_\_” heat.
* A \_\_\_\_\_\_\_ is used to measure the specific heat of a substance.



Place a hot metal (~ 75 C) in water (~25° C)

The metal \_\_\_\_\_\_\_ off (loses heat); q = \_\_\_ . The water \_\_\_\_\_\_\_ heat from the metal; q = \_\_\_

- q (metal) = q (water) \_\_\_\_\_\_\_ = \_\_\_\_\_\_\_

Thermochemical Equations

* Thermochemical equations are balanced stoichiometric chemical **equations** that include the \_\_\_\_\_\_\_ change (heat of \_\_\_\_\_\_\_ ΔH).
* Δ*Hrxn* = Δ*H*\_\_\_\_\_\_\_ – Δ*H*\_\_\_\_\_\_\_
* The \_\_\_\_\_\_\_ states (s, l, g) of reactants and products must always be included in thermochemical equations, because the states of reactants and products influence the overall amount of energy exchanged.
* In any chemical equation, the heat of reaction (∆H) is based on the enthalpy difference of the products and reactants. It is written as either a reactant or a product, depending on whether the reaction is \_\_\_\_\_\_\_ (+∆H) or \_\_\_\_\_\_\_ (-∆H).

In \_\_\_\_\_\_\_ processes, the chemical potential energy of the reactants is \_\_\_\_\_\_\_ than the chemical potential energy of the products. This is a “\_\_\_\_\_\_\_” reaction.

Because heat energy (enthalpy) is given off or \_\_\_\_\_\_\_. Therefore, the enthalpy change can be considered a \_\_\_\_\_\_\_.

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Heat (Enthalpy) of Reaction

* Heat of \_\_\_\_\_\_\_ refers to \_\_\_\_\_\_\_ Change (∆H), which is the Enthalpy or heat content of a system.
* For any reaction carried out directly at constant pressure, the heat flow (∆H = mc∆T) is exactly equal to the difference between the enthalpy of the products and the enthalpy of the reactants.

∆H = ∑[(∆H\_\_\_\_\_\_\_) × (\_\_\_\_\_\_\_ products)] – ∑[(Δ*H*\_\_\_\_\_\_\_) × (\_\_\_\_\_\_\_ reactants)]

Moles represents the \_\_\_\_\_\_\_ in a chemical equation.

Calculating Enthalpy of Reaction

Iron(III) oxide (Fe3O4(*s*), Δ*Hf* = –1117.1 kJ/mol) can react with aluminum powder to form aluminum oxide (Al2O3(*s*), Δ*Hf* = –1669.8 kJ/mol) and iron according to the equation below:

3Fe3O4(*s*) + 8Al(*s*) → 4Al2O3(*s*) + 9Fe(*s*)

What is Δ*Hrxn* for this reaction?

∆H = ∑[(∆Hproducts) × (molproducts)] – ∑[(Δ*H*reactants) × (molreactants)]

Δ*H*f (prod) = [\_\_\_ mol Al2O3(*s*) × Δ*H*f⁰ Al2O3(*s*)] + [\_\_\_ mol Fe(s) × Δ*H*f ⁰ Fe(s)]

= (\_\_\_ x \_\_\_\_\_\_\_ kJ) + (\_\_\_ x 0.00 KJ) = \_\_\_\_\_\_\_ kJ

Δ*H*f (react) = [\_\_\_ mol Fe3O4(s) × Δ*H*f⁰ Fe3O4 (s)] + [\_\_\_ mol Al × Δ*H*f⁰ Al(s)]

= (\_\_\_ x \_\_\_\_\_\_\_ kJ) + (\_\_\_ x 0.00 kJ) = \_\_\_\_\_\_\_ kJ

Δ*H*° = ∑Δ*H*f ⁰(products) – ∑Δ*H*f ⁰(reactants)

Δ*H*° = \_\_\_\_\_\_\_ kJ – (\_\_\_\_\_\_\_ kJ)

Δ*H*° = \_\_\_\_\_\_\_ kJ \_\_\_\_\_\_\_

Standard Heat of \_\_\_\_\_\_\_ Δ*H*f*⁰*

* Where do the values for ΔH originate?
* The standard heat of formation (Δ*H* f*⁰*) of a compound is the change in enthalpy that accompanies the \_\_\_\_\_\_\_ of one mole of a compound from its \_\_\_\_\_\_\_ with all substances in their standard states, those being the stable forms of the substances at 25 ⁰C and 101.3 kPa.
* The Δ*H*f⁰ of all \_\_\_\_\_\_\_ elements (including the “Professor \_\_\_\_\_\_\_” diatomic molecules is arbitrarily set at \_\_\_\_\_\_\_.
* Thus, Δ*H*f*⁰* = \_\_\_ for all elements and the diatomic molecules:

Enthalpy of Formation

* Compounds with a large \_\_\_\_\_\_\_ enthalpy of formation are very \_\_\_\_\_\_\_.
* Example: The ΔHf⁰ of CO2 is –393.5 kJ/mol of gas produced. This indicates that carbon dioxide is more stable than the elements from which it was formed.
* Compounds with \_\_\_\_\_\_\_ values of enthalpies of formation are typically \_\_\_\_\_\_\_.
* Example: Hydrogen iodide, HI, has a ΔHf⁰ of +26.5 kJ/mol. It decomposes at room temperature into violet iodine vapor, I2, and hydrogen, H2.

Standard Heats of Formation ΔH⁰

Δ*H*° = ∑Δ*H*f°\_\_\_\_\_\_\_ – ∑Δ*H*f°\_\_\_\_\_\_\_

Heat of \_\_\_\_\_\_\_

* A common ΔH, heat of reaction, is the heat of combustion, relating to the complete burning of one mole of a substance.
* All are \_\_\_\_\_\_\_.

\_\_\_\_\_\_\_’s Law

* The heat of reaction for many chemical processes cannot be \_\_\_\_\_\_\_ directly. (*Theoretically, diamond can be made into graphite.*)
* The heat released or absorbed in a chemical process is the \_\_\_\_\_\_\_ whether the process takes place in one or in several \_\_\_\_\_\_\_.
* Hess's lawshows that energy changes are “\_\_\_\_\_\_\_” functions, meaning that the amount of energy depends only on the states of the reactants and the state of the products, but not on the intermediate steps.
* “States” here refer to \_\_\_\_\_\_\_, \_\_\_\_\_\_\_, or \_\_\_\_\_\_\_.
* If a reaction can take place by more than \_\_\_\_\_\_\_ pathway and the initial and final conditions are the same, the total enthalpy change (∆Hx) is equal. Therefore, \_\_\_\_\_\_\_ the heats of reaction for the individual reactions (Intermediate pathway 🡪 ∆H1 + ∆H2).



* The overall enthalpy change in a reaction is equal to the \_\_\_\_\_\_\_ of enthalpy changes for the \_\_\_\_\_\_\_ steps in the process.

Example: Tin and Chlorine synthesize Stannic Chloride

* SnCl4 (l) is not produced directly in nature. Hess’s law allows you to determine the heat of reaction indirectly, by using the known heats of reaction of 2 or more thermochemical equations.

Step 1: Sn (s) + Cl2 (g)  ⇌ SnCl2 (s) + 83.6 kcal

Step 2: SnCl2 (s) + Cl2 (g)  ⇌ SnCl4 (l) + 46.7 kcal

Determine the OVERALL Reaction

Step 1: Sn (s) + Cl2 (g)  ⇌ ~~SnCl~~~~2~~ (s) + 83.6 kcal

Step 2: ~~SnCl~~~~2~~ (s) + Cl2 (g)  ⇌ SnCl4 (l) + 46.7 kcal

∆H rxn = ∆H1 + ∆H2 = \_\_\_\_\_\_\_ kcal + \_\_\_\_\_\_\_ kcal = \_\_\_\_\_\_\_ kcal

OVERALL \_\_\_ (s) + 2 \_\_\_2 (g)  ⇌ \_\_\_\_\_\_4 (l) + \_\_\_\_\_ kcal

\_\_\_\_\_\_\_ Reaction