**Today’s task:** Identify exothermic and endothermic reactions from their chemical equations. Calculate the heat of reaction when given the amount of reactants used.

## Part I: Identify reactions

We can add heat into a chemical equation as a reactant or a product. ALL REACTIONS REQUIRE ENERGY in order to occur, but some require lots and some only a little.

Exothermic reactions release more energy than they require.

* Heat is a product.
* Example:

CaO (s) + H2O (l) 🡪 Ca(OH)2 (s)  + **heat**

Endothermic reactions require more energy than they release.

* Heat is a reactant.
* Example: Decomposition reactions often are endothermic

2NaHCO3 (s) + **heat** 🡪 Na2CO3 + H2O (l) + CO2 (g)

Problems: Identify the following reactions as endothermic or exothermic.

1. C2H6O + 3O2 🡪 2CO2 + 3H2O + 1368 kJ \_\_\_\_\_exo\_\_\_\_\_\_\_\_\_\_\_\_
2. 2Al2O3 + 3335 kJ 🡪 4Al + 3O2 \_\_\_\_\_endo\_\_\_\_\_\_\_\_\_\_\_\_
3. 2Mg + O2 🡪 MgO + 1204 kJ \_\_\_\_\_exo\_\_\_\_\_\_\_\_\_\_\_\_
4. PCl5 + 156.5 kJ 🡪 PCl3 + Cl2 \_\_\_\_\_endo\_\_\_\_\_\_\_\_\_\_\_\_
5. Given the reactions above, what units do we use to measure heat?

Kilojoules (kJ)

## Part II: Heat of reaction

We can determine the heat given off or taken in by a reaction using calorimetry (a specific lab technique). The energy lost by a system or gained by a system during a chemical reaction is called the **enthalpy of reaction, H,** and the change in enthalpy (comparing the before and after) is usually written as **∆H**.

Exothermic reactions will have a NEGATIVE **∆H** value.

CaO (s) + H2O (l) 🡪 Ca(OH)2 (s)  + **65.2 kJ** can also be written as:

CaO (s) + H2O (l) 🡪 Ca(OH)2 (s)  **∆H = - 65.2 kJ**

Endothermic reactions will have a POSITIVE **∆H** value.

2NaHCO3 (s) + **85 kJ** 🡪 Na2CO3 + H2O (l) + CO2 (g) can also be written as:

2NaHCO3 (s) 🡪 Na2CO3 + H2O (l) + CO2 (g) **∆H = + 85 kJ**

Problems: Determine the **∆H** value for each reaction. Should the sign for this be + or -- ?

1. C2H6O + 3O2 🡪 2CO2 + 3H2O + 1368 kJ \_\_\_∆H = - 1368 kJ \_\_\_
2. 2Al2O3 + 3335 kJ 🡪 4Al + 3O2 \_\_∆H = + 3335 kJ \_\_\_\_
3. 2Mg + O2 🡪 MgO + 1204 kJ \_\_\_∆H = - 1204 kJ \_\_\_
4. PCl5 + 156.5 kJ 🡪 PCl3 + Cl2 \_\_\_∆H = + 156.5 kJ \_\_\_

## Part III: Heat of reaction calculations

We can determine the heat given off or taken in by a reaction when given the specific amount of reactants and the **∆H** value using stoichiometry. This can be done by including **∆H as a mole ratio from your balanced chemical equation.**

#### Example problems: For the following reaction:

CaO (s) + H2O (l) 🡪 Ca(OH)2 (s)  **∆H = - 65.2 kJ**

1. How much heat is produced in this reaction if you use 1 mol CaO and excess water?
2. How much heat is produced in this reaction if you use 2.5 mol CaO and excess water?
3. How much heat is produced in this reaction if you use 30 g CaO and excess water?

#### Example problems: For the following reaction:

2NaHCO3 (s) 🡪 Na2CO3 + H2O (l) + CO2 (g) **∆H = + 85 kJ**

1. What is the change in enthalpy when 2.5 mol NaHCO3 decomposes?

Therefore, ∆Hrxn = +106 kJ (i.e. 106 kJ required to decompose 2.5 mol NaHCO3)

1. What is the change in enthalpy when 30 g NaHCO3 decomposes?

Therefore, ∆Hrxn = +15 kJ (i.e. 15 kJ required to decompose 30g NaHCO3)

Problems: Determine the **∆H** value for each reaction assuming that you begin with 30 g of C2H6O, Al2O3, Mg or PCl5 in each reaction.

1. C2H6O + 3O2 🡪 2CO2 + 3H2O + 1368 kJ \_\_∆Hrxn = -890 kJ \_\_\_\_\_

Therefore, ∆Hrxn = -890 kJ (i.e. 890 kJ will be released when 30g is burned).

1. 2Al2O3 + 3335 kJ 🡪 4Al + 3O2 \_\_∆Hrxn = +491 kJ \_

Therefore, ∆Hrxn = +491 kJ (i.e. 491 kJ will be required to decompose 30g Al2O3).

1. 2Mg + O2 🡪 MgO + 1204 kJ \_\_∆Hrxn = -742 kJ \_\_\_

Therefore, ∆Hrxn = -742 kJ (i.e. 742 kJ will be released when 30g is burned).

1. PCl5 + 156.5 kJ 🡪 PCl3 + Cl2 \_\_\_,∆Hrxn = +22.5 kJ \_\_\_

Therefore, ∆Hrxn = +22.5 kJ (i.e. 22.5 kJ will be required to decompose 30g PCl5).