**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?

$$1mol CaO×\frac{-65.2 kJ}{1mol CaO}=65.2 kJ released$$

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

$$2.5mol CaO×\frac{-65.2 kJ}{1mol CaO}=163 kJ released$$

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

$$30 g CaO×\frac{1 mol CaO}{56.1g CaO}×\frac{-65.2 kJ}{1mol CaO}=34.9 kJ released$$

#### 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?

$$2.5mol NaHCO3×\frac{85 kJ}{2 mol NaHCO3}=106 kJ$$

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?

$$30 g NaHCO3×\frac{1 mol NaHCO3}{84 g NaHCO3}×\frac{85 kJ}{2 mol NaHCO3}=15 kJ$$

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 \_\_\_\_\_

$$30 g C\_{2}H\_{6}O×\frac{1 mol C\_{2}H\_{6}O}{46.1 g C\_{2}H\_{6}O}×\frac{-1368 kJ}{1 mol C\_{2}H\_{6}O}=-890 kJ$$

Therefore, ∆Hrxn = -890 kJ (i.e. 890 kJ will be released when 30g $C\_{2}H\_{6}O$ is burned).

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

$$30 g Al\_{2}O\_{3}×\frac{1 mol Al\_{2}O\_{3}}{101.96 g Al\_{2}O\_{3}}×\frac{3335 kJ}{2 mol Al\_{2}O\_{3}}=+491 kJ$$

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

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

$$30 g Mg×\frac{1 mol Mg}{24.31 g Mg}×\frac{-1204 kJ}{2 mol Mg}=-742 kJ$$

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

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

$$30 g PCl\_{5}×\frac{1 mol PCl\_{5}}{208.24 g PCl\_{5}}×\frac{156.5 kJ}{1 mol PCl\_{5}}=+22.5 kJ$$

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