

## UNIT 3. Chemical Reactions

### Hess's Law

1. **Hess's law** states that the enthalpy change for a chemical reaction is independent of the route taken.
2. Enthalpy changes can be calculated by application of Hess's law

## Hess's Law

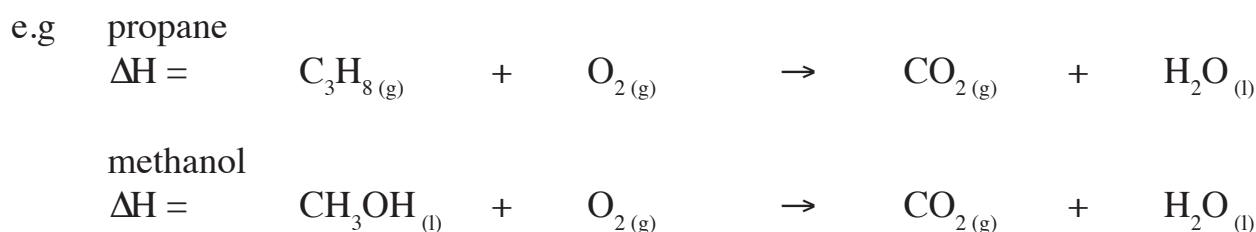
### Enthalpy Revision

This activity revises the enthalpies of **combustion**, **solution**, and **neutralisation**.

Any **che** process can have an **enth of reaction**, but certain types of reactions are deemed important enough to be **na** and defined carefully. **Val** for many of these processes are to be found in *Data Books*.

Along with the written definition, it is important that **bal equations** can also be written that describe *exactly* the reaction to which the *numerical* value of  $\Delta H$  corresponds.

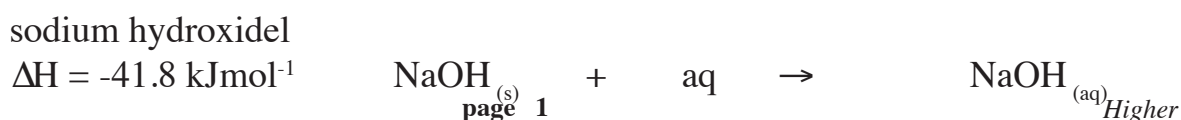
The **enthalpy of combustion** of a substance is the enthalpy change when **one mole** of the substance **burns completely** in oxygen



Notice that for the **equ** to match the  $\Delta H$  value in the Data Book it must be **bal** using 'fractions' of moles so that the **qua** of fuel remains as **one mole**.

Notice that **enth of comb** are **always exo**, ie. their  $\Delta H$  values are **always neg**.

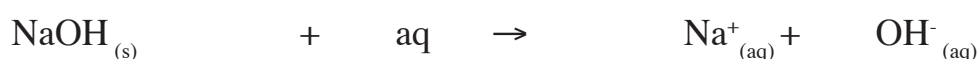
The **enthalpy of solution** of a substance is the enthalpy change when **one mole** of the substance **dissolves completely** in water



Notice that there is no *qua* for the *wa* involved; this is why *aq* is used rather than the *for*  $\text{H}_2\text{O}$ , which would imply *one mole* of water.

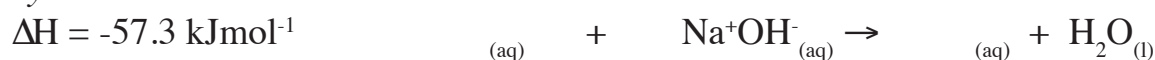
Notice that *enth of sol* can be either *exo*,  $\Delta\text{H}$  *neg*, or *endo*,  $\Delta\text{H}$  *pos*.

Once *diss* the *ionic net*, *crystal lattice*, is completely broken so the *individual ions* are now only *bon* to *wa* molecules. For this reason, these *equ* are more often written as:-

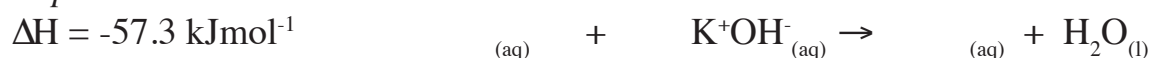


The *enthalpy of neutralisation* of an acid is the enthalpy change when the acid is neutralised to form *one mole of water*.

e.g. *hydrochloric acid* / NaOH



*sulphuric acid* / KOH



Notice, again, that there is often a need to use '*fra*' of *mo* in order to *bal* these equations in terms of *one mole of wa*, as defined.

Notice that *enth of neut* are always *exo*,  $\Delta\text{H}$  *neg*, and that the value is the *sa* for many combinations of *ac* / *alk*.

This is because the actual reaction is *iden* in most cases;



Later in the course, you will be introduced to *weak acids* like *ethanoic acid*,  $\text{CH}_3\text{COOH}$ , and *weak alkalis* like *ammonium hydroxide*,  $\text{NH}_4\text{OH}$ . Acids like these have *different* enthalpies of neutralisation.

For any combination of *strong acid* / *strong alkali*, however, the enthalpy of neutralisation is always very close to  $-57.3 \text{ kJmol}^{-1}$ .

## Specific Heat Capacity Revision

$\Delta H$  = energy gained or lost by the water

$c$  = specific heat capacity of the water

$m$  = mass of the water

$\Delta T$  = rise or fall in temperature

$$\Delta H = c \times m \times \Delta T$$

This activity looks at how the **specific heat capacity** of water can be used to experimentally determine the enthalpy of a reaction.

During **chem** reactions **en** is **rel** to the **surr**, **thermic**, or **tak in** from the **surr**, **thermic**.

Often, the **surr** are the **wa** that the chemicals are **diss** in.

We know exactly **how much energy** it takes to heat, or cool down, **1kg** of water by exactly **1°C**. This is the **specific heat capacity**,  $c = 4.18 \text{ kJ kg}^{-1} \text{ °C}^{-1}$

## Hess's Law PPA

PPA

This activity looks at how Hess's law can be proven by experiment

Hess's Law states:

'The enthalpy change of a chemical reaction depends only on the chemical nature and physical states of the reactants and products and is independent of any intermediate

\* State the aim of the experiment

### Equations

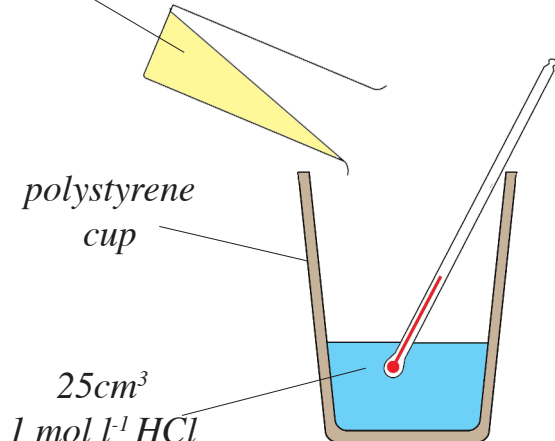
**Pot hyd** (KOH) can be converted into **pot chl** (KCl) by two different routes

\* Use equations to describe the two routes whereby you converted solid potassium hydroxide into potassium chloride and label them with appropriate  $\Delta H$  values

**Procedure**

1.2 g of potassium hydroxide

*Route 1*

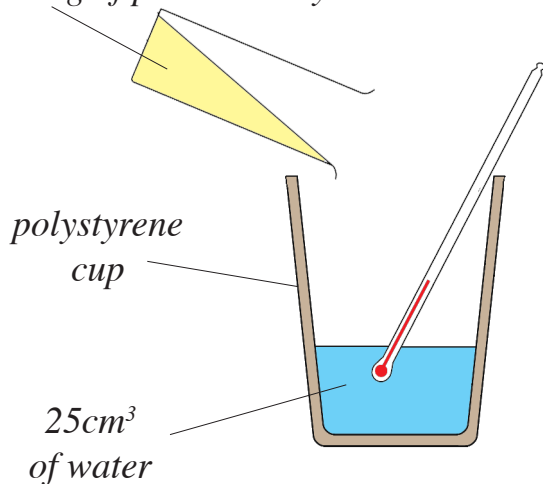


The *initial temp* of the *ac* was measured.

The was added and *sti* with the *therm*. The *highest temp* reached was recorded

1.2 g of potassium hydroxide

*Route 2 - step a*

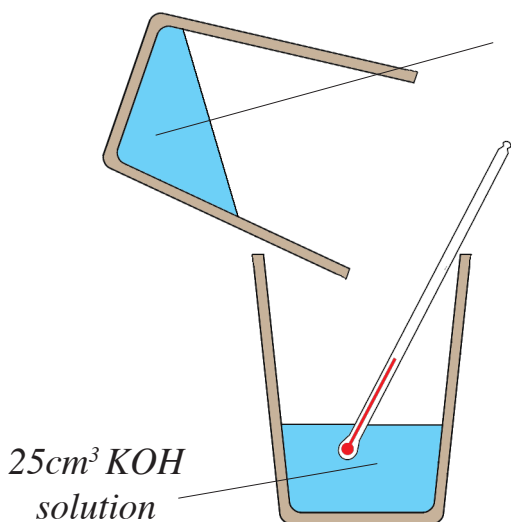


The *initial temp* of the *wa* was measured.

The was added and *sti* with the *therm*. The *highest temp* reached was recorded

25cm<sup>3</sup>  
1 mol l<sup>-1</sup> HCl

*Route 2 - step b*



The *initial temp* of the *ac* and the KOH *sol* were measured and an *av starting temp* determined.

The HCl was added and *sti* with the *therm*. The *highest temp* reached was recorded

## ***Results***

\* *Present your results in an appropriate manner.*

## ***Calculations/Conclusion***

\* *carry out your calculations to show the confirmation of Hess's Law*

## Hess's Law Calculations

### Using Diagrams

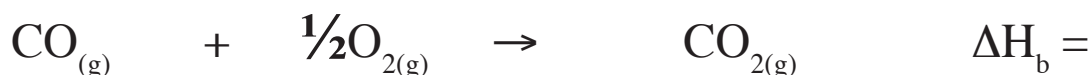
*This activity shows how enthalpies of reactions can be calculated using other experimentally determined enthalpies.*

Only certain enthalpies can be measured directly. Hess's Law allows us to calculate enthalpy changes that are very difficult or even impossible to measure.

For example, carbon burns to produce carbon monoxide but it is impossible to burn carbon in such a way that the **only** product is carbon monoxide. ie. we cannot measure the  $\Delta H$  for this reaction:



We can, however, measure the enthalpies of these two reactions (values in Data Book).



There are two routes that we can use to convert  $\text{C}_{(s)} \rightarrow \text{CO}_{2(g)}$

A diagram can be drawn to show these two routes:

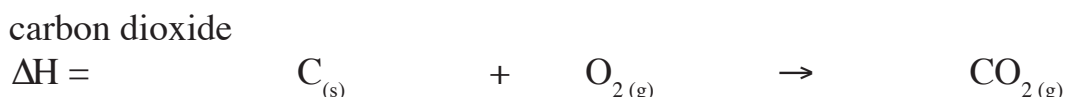
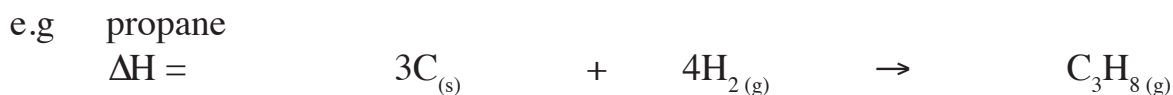
Using the diagram, and Hess's Law, it is clear that:

$$\Delta H_a = \Delta H_c + \Delta H_b$$

Putting in the values we know, will allow us to calculate  $\Delta H_c$

An extremely useful enthalpy to know is called the enthalpy of formation of a substance: by comparing enthalpies of formation chemists learn a lot about relative stabilities and the energy changes needed to change one substance into another,

*The **enthalpy of formation** of a substance is the enthalpy change when **one mole** of the substance **is formed** from its elements in their natural state*



Most of these enthalpies, particularly of organic molecules, cannot be measured experimentally; either the reactions are impossible or there will be a mixture of products.

Many of the enthalpies can be measured. Notice that the enthalpy of formation of carbon dioxide is also the enthalpy of combustion of carbon.

Again a diagram can be constructed that links a reaction we cannot measure;



to reactions that can be measured;

