Name: ______________________ Date: __________ Period: ______

SL IB Chem 1 Thermodynamics - Guided Notes

Slide 3: Measuring Enthalpy

- **Think of the burning food as your ____________.**
- **The ____________ change in the water is related to the energy given off.**

Slide 4: Enthalpy Change

- We calculate enthalpy change directly from the ____________ change of a known mass of water.

Slide 5: Enthalpy Change

Slide 6: Specific Heat Capacity

Slide 7: Measuring enthalpy changes

- **____________ state is the state at RTP (SATP):**
- **100 kPa = 100 000 Pa**
- **25ºC = 298 K**
- **Standard ____________ and specific heat capacity of water are in your data books.**
- **____________ of water = 1g/cm^3**
- **1g water = 1cm^3**
- **(Be ___________ - not all liquids are the same)**

Slide 8: Example

- The student carried out a laboratory experiment to determine the enthalpy change when a sample of butan-1-ol was burned. The student found that the temperature of 175 g of water increased by 8.0 ºC when 5.00 x 10^{-3} mol of pure butan-1-ol was burned in air and the heat produced was used to warm the water. Use the student's results to calculate a value, in kJ mol^{-1}, for the enthalpy change when one mole of butan-1-ol is burned. (The specific heat capacity of water is 4.18 J K^{-1} g^{-1})
- ***What information from here do we need???***

Slide 9: Example

- The student carried out a laboratory experiment to determine the enthalpy change when a sample of butan-1-ol was burned. The student found that the temperature of **175 g of water** increased by **8.0 °C** when 5.00 x 10^{-3} mol of pure butan-1-ol was burned in air and the heat produced was used to warm the water. Use the student's results to calculate a value, in kJ mol^{-1}, for the enthalpy change when one mole of butan-1-ol is burned. (The specific heat capacity of water is **4.18 J K^{-1} g^{-1}**)
- *m = 175 g*
- *c = 4.18 $JK^{-1}g^{-1}$*
- *ΔT = 8 K*

Slide 10: Example

- 5.852 / (5.00x10^{-3})
- When doing a calorimetry experiment, the heat released by an **exothermic reaction is absorbed by water.**
- $\Delta H_{rxn} = -\Delta H_{water}$

Slide 11: Enthalpy change in solution

Slide 12: Enthalpy change in solution

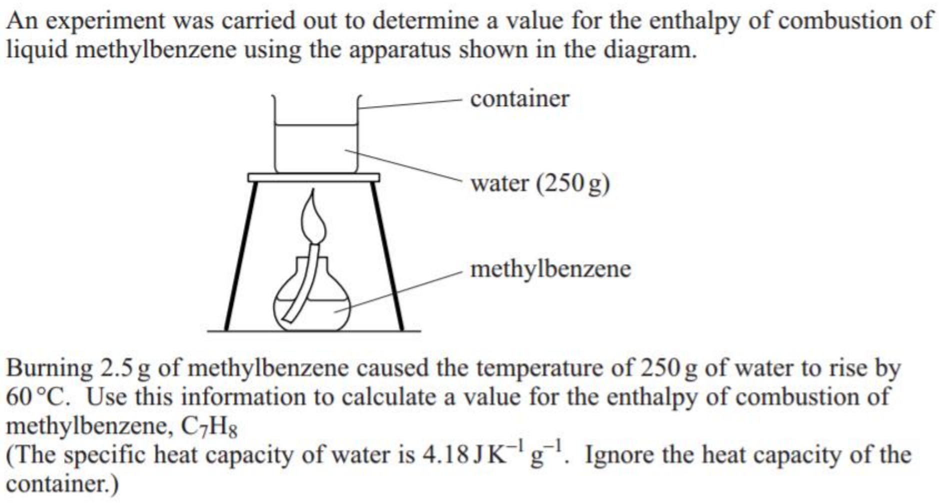
- 50.0 cm^3 of 1.0 mol dm^{-3} $NaOH_{(aq)}$ is added to 50.0 cm^3 of 1.0 mol dm^{-3} $HCl_{(aq)}$ in a calorimeter. The temperature of the reaction mixture rises from 18.0°C to 24.6°C. The specific heat capacity of water is 4.18 J K^{-1} g^{-1}. Calculate the standard molar enthalpy of ____________.
- *1 cm^3 water = 1g*
- m = 50 + 50 = 100 g
- c = 4.18 J K^{-1} g^{-1}.
- ΔT = 24.6 – 18.0 = 6.6 K
- Q = 100 x 4.18 x 6.6 = 2758.8 J

- Q = 2.7588kJ
- n= CxV = 0.05 x 1
- = 0.05 mol
- ΔH_n^0 = – (q ÷ n)
- = – 2.7588 / 0.05
- = – 55.18 kJ mol^{-1}

Slide 13: Theory of knowledge

- What criteria do we use in judging discrepancies between experimental and theoretical values? Which ways of knowing do we use when assessing experimental limitations and theoretical assumptions?

Slide 14: Exam Question



An experiment was carried out to determine a value for the enthalpy of combustion of liquid methylbenzene using the apparatus shown in the diagram.

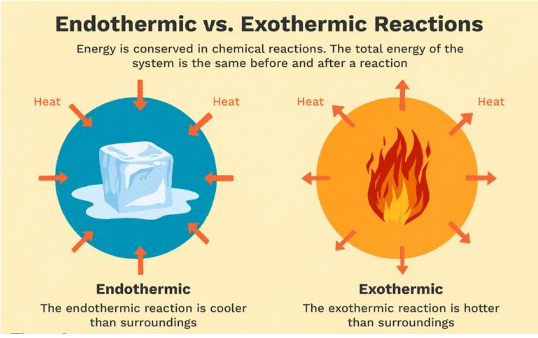
Burning 2.5 g of methylbenzene caused the temperature of 250 g of water to rise by 60 °C. Use this information to calculate a value for the enthalpy of combustion of methylbenzene, C_7H_8
(The specific heat capacity of water is 4.18 J K^{-1} g^{-1}. Ignore the heat capacity of the container.)

Slide 15: List all of the problems with this method.

- **Calorimetric Methods**

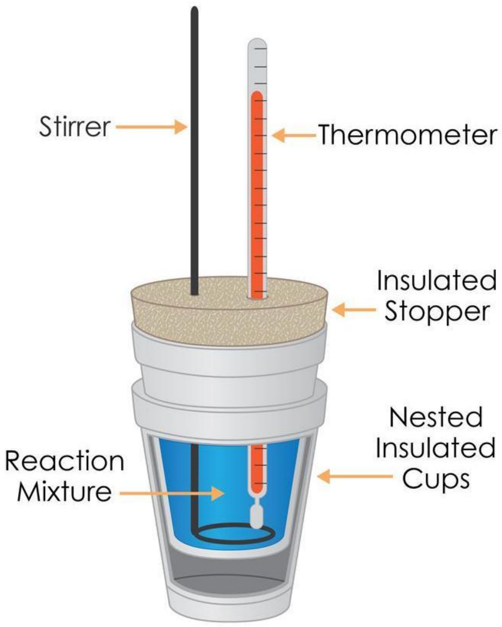
Slide 16: Experiments investigating endothermic and exothermic processes could be run numerous times to compare reliability of repetitive data and compare to theoretical values.

- **Reliable data?**



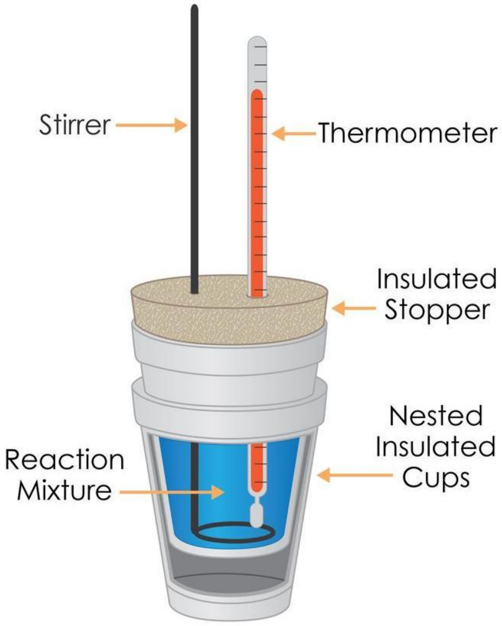
Slide 17: Heat Loss

- The _____________ beaker method helps us to reduce heat loss because polystyrene is a good insulator.
- (Only for _____________ in solution)



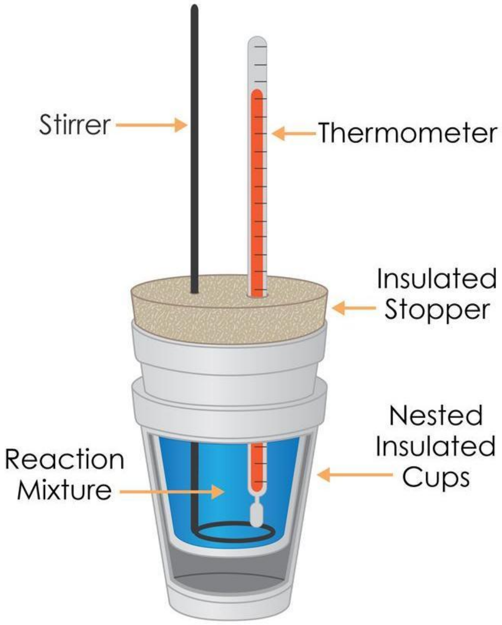
Slide 18: Heat Loss

- Can heat loss still occur?
- How will this affect your measurement of enthalpy?



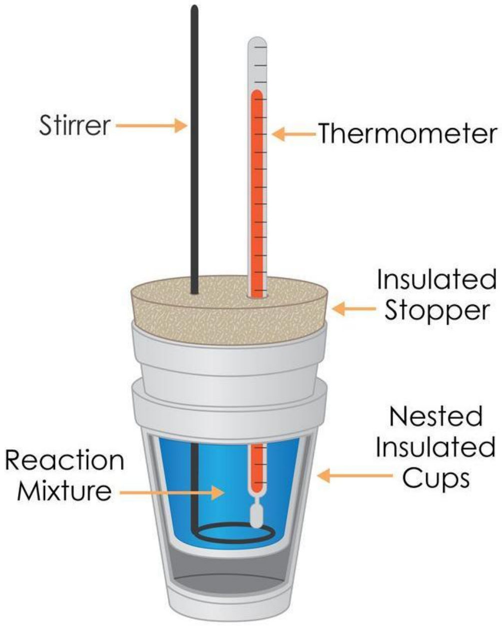
Slide 19: Allowing for heat loss

- The way in which we ____________ data can affect our final results
- Use a **cooling curve** to find the **best ____________ of ΔT**



Slide 20: Uncertainty

- **How many uncertainties can you find for a calorimetry experiment?**
- **Temperature measurements**
- **Volume/mass of water**
- **Volume/mass of reactants**
- **How do we quantify these uncertainties?**



Slide 21: Uncertainty

- Consider a _____________ experiment measuring the enthalpy change when 1 mol dm^{-3} NaOH is mixed with 1 mol dm^{-3} HCl (slide 11).
- _____________: uncertainty is half the smallest division (analog) or 0.1 °C (digital).
- If the _____________ is measured as 25.6 C, it is written as T = 25.6 ± 0.1 °C, meaning it could be between 25.5 and 25.7 °C. ± 0.1 °C is the absolute uncertainty.
- Change this into a percentage _____________:
- 0.1/25.6 x 100 = 0.39%

Slide 22: Error Propagation

- **To find ΔT we take 2 _____________: initial and final. Each measurement has an uncertainty.**
- **ΔT = 25.6-19.0 = 6.6 ˚C**
- **When we add/subtract measurements, we ADD the absolute _____________.**
- **ΔT = 6.6 ± 0.2˚C = 6.6 ˚C ± 3.03%**
- **When we multiply/divide measurements, we ADD the percentage _____________.**

Slide 23: Practice

- Calculate the uncertainty when making a 1 mol dm^{-3} solution of NaOH.
- **A 2 d.p. mass balance was used to weigh 10.00g of solid sodium hydroxide.**
- **In making the standard solution, a 250 cm^3 volumetric flask (±0.12 cm^3) was used.**
- **Assume M_r(NaOH)=39.997 g mol^{-1} with no error.**
- **0.01/10.00 = 0.10% uncertainty**
- **0.12/250.00 = 0.048% uncertainty**
- **[NaOH] = 1.00 mol dm^{-3} ± 0.15%**

Slide 24: Practice

- ΔT = 6.6 ± 0.2˚C = 6.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- **50.0 cm^3 of each solution (HCl and NaOH) were added together to the calorimeter using 50.0 cm^3 graduated pipettes with an uncertainty of ±0.05 cm^3.**
- **0.10/100.00 = 0.10% uncertainty**
- **Volume: 100.0 cm^3 ± 0.10%**

Slide 25: Practice

- ΔT = 6.6 ± 0.2˚C = 6.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- Volume HCl + NaOH = 100.0 cm^{3} ± 0.10%
- n(NaOH)= 0.0500 mol ± 0.25%
- **Calculate the % uncertainty in the number of moles of base in the reaction.**
- 1.00 mol dm^{-3} x 0.0500 dm^{-3} = 0.0500 mol NaOH
- 0.15% + 0.10% = 0.25%

Slide 26: Practice

- ΔT = 7.6 ± 0.2˚C = 7.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- Volume HCl + NaOH = 100.0 cm^{3} ± 0.10%
- n(NaOH)= 0.0500 mol ± 0.25%
- mass of water = 100.0 g ± 0.10%
- **Assume density of water is 1.00 g cm^{-3}**
- **Calculate the % uncertainty in the volume of the reaction mixture and therefore the mass of water.**

Slide 27: Practice

- ΔT = 6.6 ± 0.2˚C = 6.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- Volume HCl + NaOH = 100.0 cm^{3} ± 0.10%
- n(NaOH)= 0.0500 mol ± 0.25%
- mass of water = 100.0 g ± 0.10%

- **Calculate the % uncertainty in the enthalpy change of the reaction. Assume specific heat capacity of water = 4.18 $Jg^{-1}K^{-1}$ with no error.**
- **q = mcΔT**
- **= 100.0(4.18)(6.6) = 2,758.8 J ± 3.13%**

Slide 28: Practice

- ΔT = 7.6 ± 0.2˚C = 7.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- Volume HCl + NaOH = 100.00 cm^3 ± 0.10%
- n(NaOH)= 0.0500 mol ± 0.25%
- mass of water = 100.00 g ± 0.10%
- ΔH = 2758.8 J ± 3.13%
- **Calculate the % uncertainty in the enthalpy change per mole of the reaction.**
- **2758.8 J/0.0500 mol = 55,176 J mol^{-1} ± 3.38%**

Slide 29: Practice

- ΔT = 7.6 ± 0.2˚C = 7.6 ˚C ± 3.03%
- [NaOH] = 1.00 mol dm^{-3} ± 0.15%
- Volume HCl + NaOH = 100.00 cm^3 ± 0.10%
- n(NaOH)= 0.0500 mol ± 0.25%
- mass of water = 100.00 g ± 0.10%
- ΔH = 2758.8 J ± 3.13%
- **ΔH^0 = -55.18 kJ ± 3.38%**
- **Therefore there is an uncertainty of 1.87kJ in the calculation.**

Slide 30: Practice

- Consider an experiment measuring the molar enthalpy change involved in dissolving anhydrous copper sulfate in water.
- Anhydrous copper sulfate is added to a dry calorimeter, and 25.0 cm^3 of distilled water is added. The temperature change is measured while the solution is stirred. The initial temperature was 26.1°C ± 0.1 °C and the final temperature was 57.9 °C ± 0.1 °C.
- Complete a full error analysis and calculate an absolute uncertainty for the standard molar enthalpy change for the reaction.
- Assume: 7.98 g of anhydrous copper sulphate
- 2 d.p. mass balance
- 25.0 cm^3 ± 0.06 cm^3 pipette

Slide 31: Practical

- This unit, we will do a ____________ practical.
- **You will be ____________ to**
- **plan and carry out a ____________ method**
- **evaluate sources of error in your method, apparatus and ____________ technique.**
- **carry out ____________ error analysis on your results.**

Slide 32: Extrapolating graphs

Slide 34: Energy vs. Heat

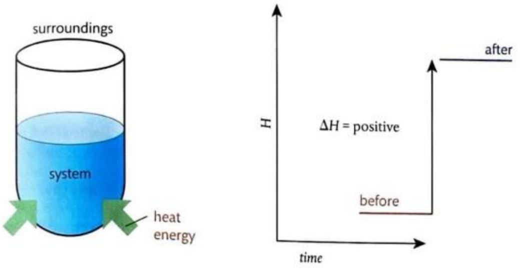
- Energy is the ability to do work and ____________:
- Heat
- Light
- Sound
- Electricity
- Chemical energy
- Heat is a mode of energy transfer as a result of ____________ differences
- Produces an increase in average kinetic energy of ____________
- Causes an increase in disorder of ____________
- **Energy ____________/released during chemical rxns**

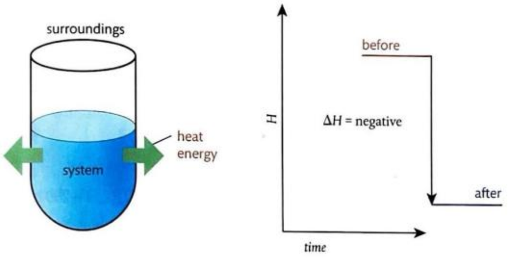
Slide 35: System vs. Surroundings

- When studying energy transfer, it is important to define the **system** and its **_____________.**
- **Area of interest**
- **_____________ else in the universe**
- Most chemical reactions happen in **open systems** (able to exchange energy and matter with _____________)
- Closed systems can exchange energy but not matter with the _____________.
- **_____________, TOTAL energy doesn't change during a reaction.**

Slide 36: Enthalpy

- The heat _____________ of a system (H)
- Changes in _____________ (ΔH) are positive when heat is added to a system
- Changes in enthalpy are negative when heat is released from the system to the _____________.





Slide 37: Endothermic vs. Exothermic Rxns

- $\Delta H < 0$
- $\Delta H > 0$

Slide 38: Standard Enthalpy Changes

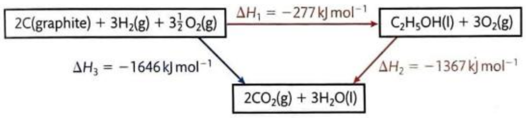
- $\Delta H^{\ominus}$: _____________ enthalpy change
- “Standard” means measured under the following _____________:
- 100kPa
- 1 mol dm^{-3} _____________ for all solutions
- All _____________ in their standard states
- Pure form of a substance under standard _____________ of 298 K (25 °C) and 100 kPa

Slide 39: Thermochemical Equations

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

- $\Delta H^{\ominus}$ is used to describe enthalpy changes in chemical ____________.
- $\Delta H^{\ominus}$= -890 kJ mol^{-1}
- $6CO_2(g) + 6H_2O\ (l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
- $\Delta H^{\ominus}$= +2802.5 kJ mol^{-1}
- $\Delta H < 0$, exothermic
- $\Delta H > 0$, endothermic

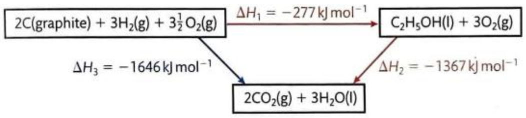
Slide 40: Hess's Law



- ____________ enthalpy changes of a reaction are not measured directly.
- The ΔH of a rxn is ____________ from the known enthalpy changes of other reactions.
- ____________ route (red arrows): elements combined to form ethanol, which is burned.
- ____________ route (blue arrow): elements burned separately

Slide 41: Hess's Law

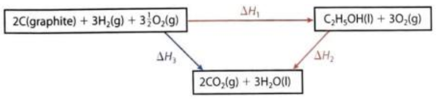
- **Clockwise Route:**



$$\Delta H_1 + \Delta H_2 = -277 + -1367 = -1644\ kJ\ mol^{-1}$$

$$\Delta H_3 = \Delta H_1 + \Delta H_2$$

Slide 42: Hess's Law



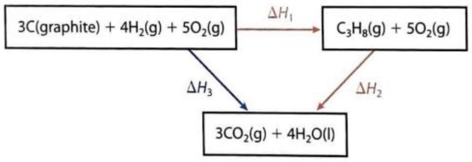
- **For the Law of ____________ of Energy to be followed, $\Delta H_1 + \Delta H_2$ must be the same as ΔH_3, since both correspond to the combustion of 2 moles of carbon and 3 moles of hydrogen.**
- **____________ of the multiple stages or steps of a reaction, the total enthalpy change for the reaction is the sum of all changes.**

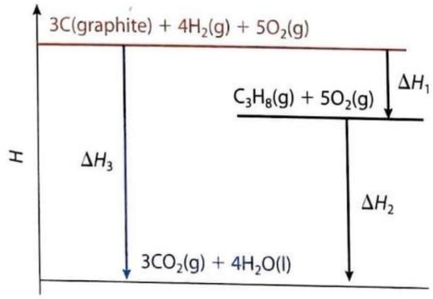
Slide 43: Using Hess's Law

- **An energy level diagram can be used to show the**

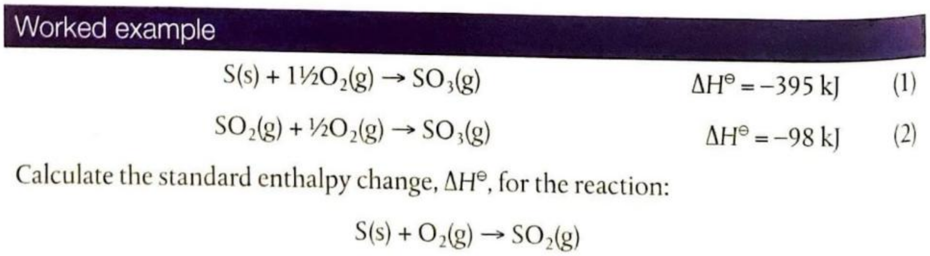
____________ between different reactions.

- **Even though ΔH_1 can't be measured directly, it can be ____________ from the ΔH_c of carbon, hydrogen, and propane.**





Slide 44: Using Hess's Law-Example



Worked example

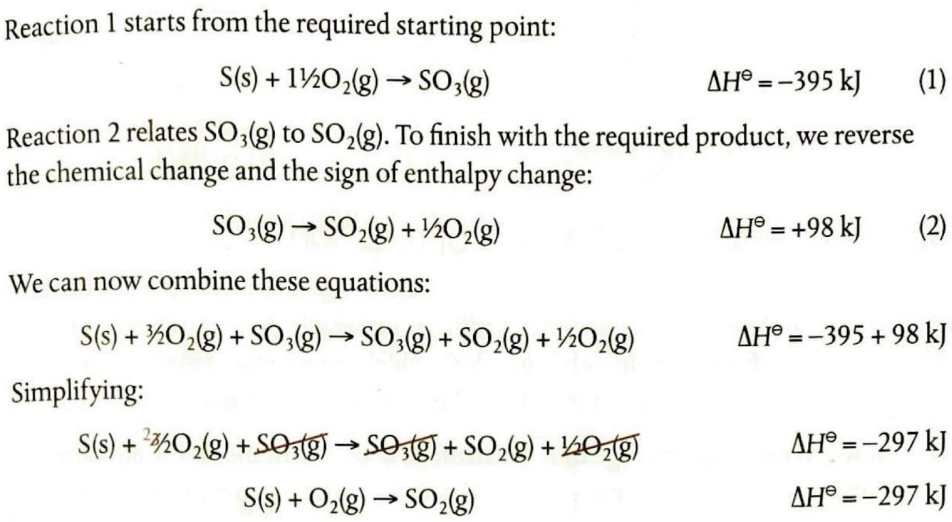
$S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$ $\Delta H^{\ominus} = -395\ kJ$ (1)

$SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$ $\Delta H^{\ominus} = -98\ kJ$ (2)

Calculate the standard enthalpy change, $\Delta H^{\ominus}$, for the reaction:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Slide 45: Using Hess's Law-Example



Reaction 1 starts from the required starting point:

$$S(s) + 1\tfrac{1}{2}O_2(g) \rightarrow SO_3(g) \qquad \Delta H^{\ominus} = -395\ \text{kJ} \qquad (1)$$

Reaction 2 relates $SO_3(g)$ to $SO_2(g)$. To finish with the required product, we reverse the chemical change and the sign of enthalpy change:

$$SO_3(g) \rightarrow SO_2(g) + \tfrac{1}{2}O_2(g) \qquad \Delta H^{\ominus} = +98\ \text{kJ} \qquad (2)$$

We can now combine these equations:

$$S(s) + \tfrac{3}{2}O_2(g) + SO_3(g) \rightarrow SO_3(g) + SO_2(g) + \tfrac{1}{2}O_2(g) \qquad \Delta H^{\ominus} = -395 + 98\ \text{kJ}$$

Simplifying:

$$S(s) + {}^{2}\cancel{3}/2\,O_2(g) + \cancel{SO_3(g)} \rightarrow \cancel{SO_3(g)} + SO_2(g) + \cancel{\tfrac{1}{2}O_2(g)} \qquad \Delta H^{\ominus} = -297\ \text{kJ}$$

$$S(s) + O_2(g) \rightarrow SO_2(g) \qquad \Delta H^{\ominus} = -297\ \text{kJ}$$

Slide 46: You Try!

- **Calculate the enthalpy change, ΔH⊖, for the reaction**
- **$C(graphite) + \tfrac{1}{2}O_2(g) \rightarrow CO(g)$**
- **From the information below:**
- **$C(graphite) + O_2(g) \rightarrow CO_2(g)$ ΔH⊖ = -394 kJ**
- **$CO(g) + \tfrac{1}{2}O_2(g) \rightarrow CO_2(g)$ ΔH⊖ = -283 kJ**
- **Calculate the enthalpy change, ΔH⊖, for the reaction**
- **$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$**
- **From the information below:**
- **$N_2(g) + O_2(g) \rightarrow 2NO(g)$ ΔH⊖ = +180.5 kJ**
- **$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ ΔH⊖ = +66.4 kJ**

Slide 47: More Hess's Law Problems

- **Calculate the enthalpy change, ΔH⊖, for the dimerization of nitrogen dioxide**

- **$2NO_2$ (g) → N_2O_4(g)**
- **From the information below:**
- **½N_2(g) + O_2(g) → NO_2(g) ΔH⊖ = +33.2kJ mol^{-1}**
- **N_2(g) + $2O_2$(g) → N_2O_4(g) ΔH⊖ = +9.16 kJ mol^{-1}**
- You can also multiply/divide standard enthalpy changes to get to the final equation.
- **2[NO_2(g) → ½N_2(g) + O_2(g)] ΔH⊖ = 2(-33.2kJ mol^{-1})**
- **N_2(g) + $2O_2$(g) → N_2O_4(g) ΔH⊖ = +9.16 kJ mol^{-1}**
- **$2NO_2$ (g) + N_2(g) + $2O_2$(g) → N_2(g) + $2O_2$(g) + N_2O_4(g) ΔH⊖ = -57.24 kJ mol^{-1}**

Slide 48: You Try!

- **The thermochemical equations for three related reactions are shown. Determine ΔH_3**
- **CO(g) + ½ O_2(g) → CO_2(g) ΔH_1= -283 kJ mol^{-1}**
- **$2H_2$(g) + O_2(g) → $2H_2O$(l) ΔH_2 = -572 kJ mol^{-1}**
- **CO_2(g) + H_2(g) → CO(g) +H_2O(l) ΔH_3 = ?**

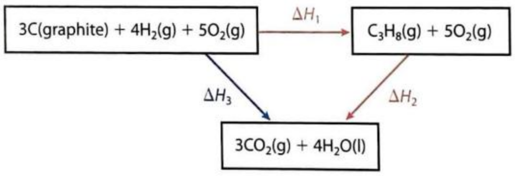
Slide 49: Enthalpy of Formation

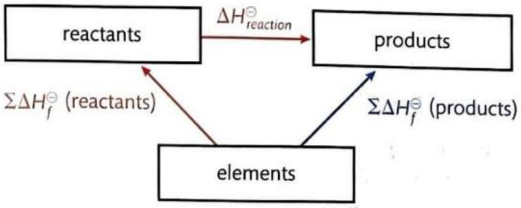
- **Standard Enthalpy of ____________** (ΔH_f⊖): the enthalpy change that occurs when **one mole** of the substance is formed from its elements in their **standard states**.
- Gives a measure of the ____________ of a substance relative to its elements
- Can be used to calculate the enthalpy change of **all reactions** (____________ or real)
- Standard Enthalpy of ____________ of Ethanol:
- 2C (____________) + $3H_2$(g) + ½O_2(g) → C_2H_5OH(l) ΔH = -278 kJ mol^{-1}
- **One mole** of product formed, so sometimes **fractional ____________** used in the equation
- **Note: ΔH_f⊖ for pure ____________ are ZERO**

Slide 50: Helpful Hints

- For ____________ reactions, **add** ΔH_1 and ΔH_2 to get ΔH_3
- For ____________ reactions, **subtract** ΔH_f(reactants) from ΔH_f(products)

- $\Delta H^{\ominus}_{rxn} = \sum \Delta H_f^{\ominus}(\text{products}) - \sum \Delta H_f^{\ominus}($____________$)$



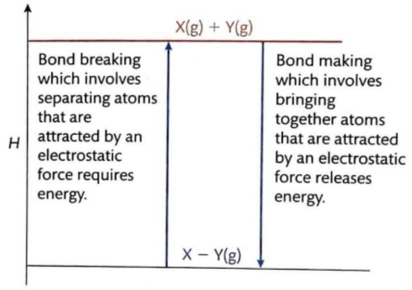


Slide 51: Example Problem

- Calculate the enthalpy change for the reaction:
- $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
- $\Delta H^{\ominus}_{rxn} = \sum \Delta H_f^{\ominus}(\text{products}) - \sum \Delta H_f^{\ominus}(\text{reactants})$
- $= 3(-394) + 4(-286) - (-105)$
- $= -2221$ kJ mol^{-1}

Slide 52: Bond Enthalpies

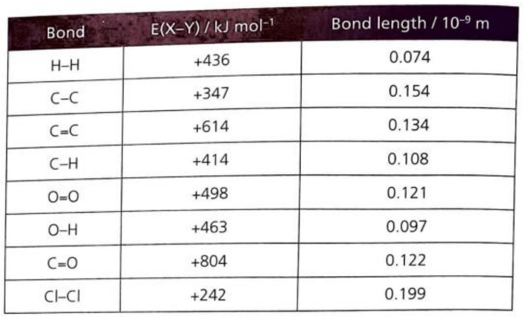
- Breaking bonds is ____________
- Making bonds is ____________
- Energy is needed to separate atoms ____________ bonded
- Energy is ____________ when bonds are formed
- Chemical reactions involve **bond making** and **bond breaking**. If more energy is released than absorbed, the reaction with be **overall** ____________
- $Cl_2(g) \rightarrow 2Cl(g)$ $\Delta H^{\ominus} = +242$ kJ mol^{-1}
- $H(g) + H(g) \rightarrow H_2(g)$ $\Delta H^{\ominus} = -436$ kJ mol^{-1}



Slide 53: Bond Enthalpies

- **Average bond enthalpy:** the energy needed to break one mole of bonds in **gaseous molecules** under standard ____________ **averaged** over **similar compounds**

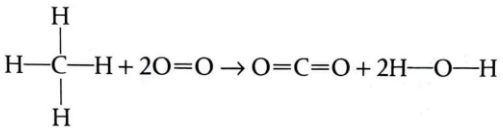
- Multiple bonds generally have **higher bond** ____________ and **shorter bond lengths** compared to single bonds



Bond	E(X–Y) / kJ mol^{-1}	Bond length / 10^{-9} m
H–H	+436	0.074
C–C	+347	0.154
C=C	+614	0.134
C–H	+414	0.108
O=O	+498	0.121
O–H	+463	0.097
C=O	+804	0.122
Cl–Cl	+242	0.199

Slide 54: Bond Enthalpies

- $\Delta H^{\ominus}_{rxn} = \sum\Delta H^{\ominus}(\text{bonds broken}) - \sum\Delta H^{\ominus}(\text{bonds formed})$
- Example: Combustion of Methane
- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
- $\Delta H^{\ominus} = +2652 - 3460$
- $= -808$ kJ mol^{-1}



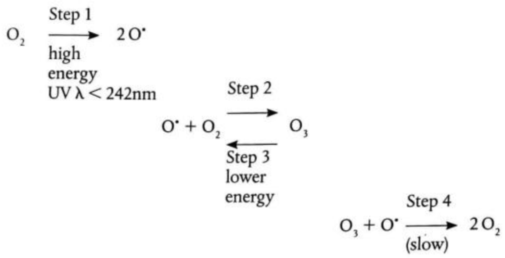
$$\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}{-}H + 2O{=}O \rightarrow O{=}C{=}O + 2H{-}O{-}H}$$

Slide 55: Ozone Depletion

- Bond order of ozone: 3/2 = 1.5
- Bond order of oxygen: 2
- Oxygen has stronger bonds than ozone, which break under lower energy/longer ____________ of light

Slide 56: The Chapman Cycle

- Ozone is ____________ at a **steady state** if the rate of formation of ozone is balanced by its rate of removal
- ____________ free radicals are formed when oxygen is broken with UV light
- Formation of ozone is ____________ because bonds are formed
- Formation of oxygen from ozone and free radicals is also ____________



Slide 57: Example Problem

- The bond energy in ozone is 363 kJ mol^{-1}. Calculate the wavelength of UV radiation needed to break the bond.
- $L \times E_{photon} = 363{,}000$ J
- $E_{photon} = 363{,}000\ J/6.02 \times 10^{23} = 6.03 \times 10^{-19}$ J
- $E = hf$,
- $c = \lambda f$ so $f = c/\lambda$
- $E = hc/\lambda$ or
- $\lambda = ch/E$
- $\lambda = 3.00 \times 10^{8} ms^{-1}(6.63\times10^{-34} Js)/6.03 \times 10^{-19}$ J
- $= 3.30 \times 10^{-7}$ m = 330 nm
- Avogadro's Number!

Slide 58: You Try!

- Use the Data Booklet to calculate the minimum wavelength of radiation needed to break the O=O double bond in O_2.

Slide 59: Additional Higher Level

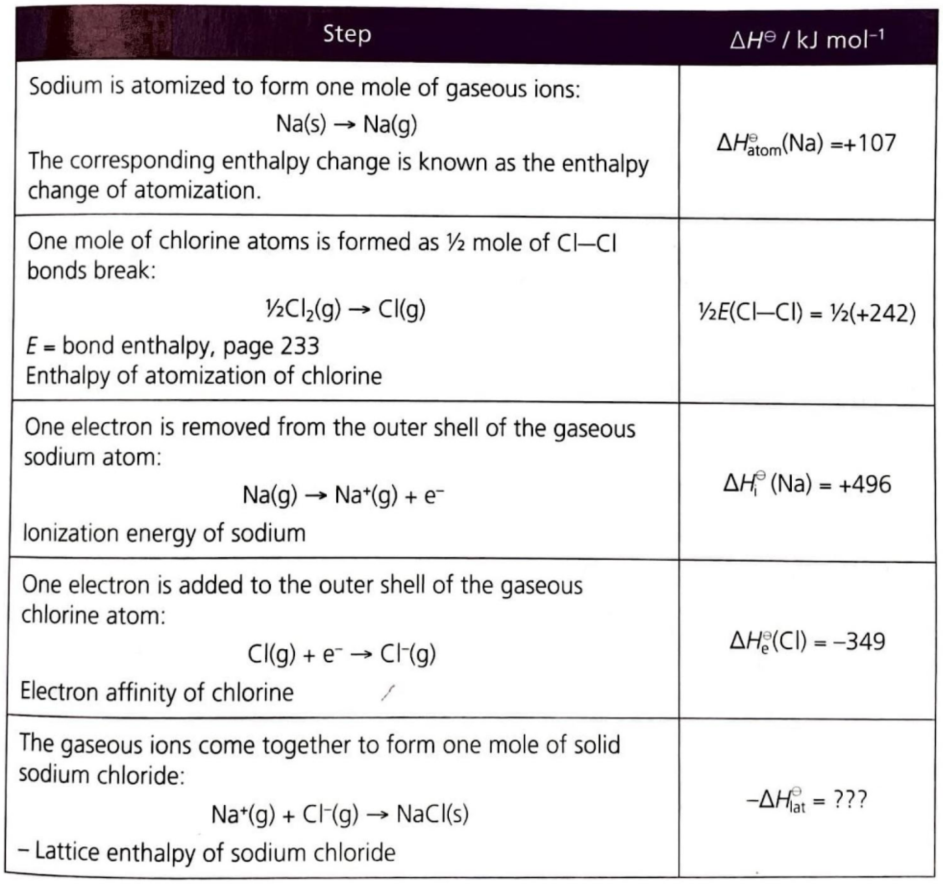
Slide 60: Some Useful Definitions

- **First _____________ Energy:** the minimum energy required to remove one mole of electrons from one mole of gaseous atoms.
- **Electron Affinity:** the enthalpy change when one mole of gaseous _____________ is added to one mole of gaseous atoms.
- **Lattice Enthalpy:** the enthalpy change that occurs when one mole of a solid ionic compound is separated into gaseous ions under standard _____________.
- $Na(g) \rightarrow Na^{+}(g) + e^{-}(g)$ $\Delta H_i^{\ominus} = +496$ kJ mol^{-1}
- $Cl(g) + e^{-}(g) \rightarrow Cl^{-}(g)$ $\Delta H_e^{\ominus} = -349$ kJ mol^{-1}
- $NaCl(s) \rightarrow Na^{+}(g) + Cl^{-}(g)$ $\Delta H_{lat}^{\ominus} = +790$ kJ mol^{-1}

Slide 61: Lattice Enthalpies

- Gaseous ions coming together to form an ionic lattice is an **exothermic process**, but lattice enthalpies express the reverse _____________ process (breaking apart a solid ionic lattice into its gaseous ions) .
- Experimental lattice enthalpies can't be determined _____________.
- The **Born-Haber Cycle** (based on Hess's Law) is used.
- Shows the enthalpy change for the _____________ of a solid ion from its gaseous elements through several steps. Final enthalpy shown will be - $\Delta H_{lat}^{\ominus}$

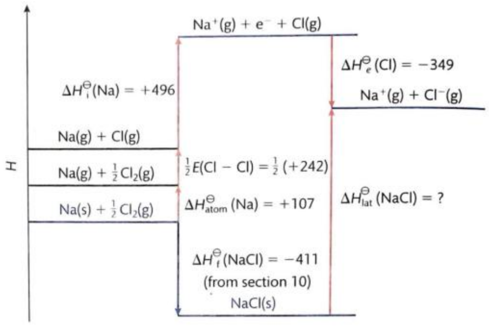
Slide 62: Born-Haber Cycle



Step	$\Delta H^{\ominus}$ / kJ mol^{-1}
Sodium is atomized to form one mole of gaseous ions: $Na(s) \rightarrow Na(g)$ The corresponding enthalpy change is known as the enthalpy change of atomization.	$\Delta H_{atom}^{\ominus}(Na) = +107$
One mole of chlorine atoms is formed as ½ mole of Cl—Cl bonds break: $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$ E = bond enthalpy, page 233 Enthalpy of atomization of chlorine	$\frac{1}{2}E(Cl{-}Cl) = \frac{1}{2}(+242)$
One electron is removed from the outer shell of the gaseous sodium atom: $Na(g) \rightarrow Na^+(g) + e^-$ Ionization energy of sodium	$\Delta H_i^{\ominus}(Na) = +496$
One electron is added to the outer shell of the gaseous chlorine atom: $Cl(g) + e^- \rightarrow Cl^-(g)$ Electron affinity of chlorine	$\Delta H_e^{\ominus}(Cl) = -349$
The gaseous ions come together to form one mole of solid sodium chloride: $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$ – Lattice enthalpy of sodium chloride	$-\Delta H_{lat}^{\ominus}$ = ???

Slide 63: Born-Haber Cycle

- _____________ of one mole of NaCl from its elements in their **standard states**
- The _____________ of one mole of NaCl from its gaseous ions (shown as -$\Delta H_{lat}^{\ominus}$)
- $\Delta H_f^{\ominus}(NaCl) = \Delta H_{atom}^{\ominus}(Na) + \Delta H_i^{\ominus}(Na) + \frac{1}{2}E(Cl\text{-}Cl) + \Delta H_e^{\ominus}(Cl) - \Delta H_{lat}^{\ominus}(NaCl)$
- $-411 = +107 + 496 + \frac{1}{2}(+242) - 349 - \Delta H_{lat}^{\ominus}(NaCl)$
- $\Delta H_{lat}^{\ominus}(NaCl) = +786$ kJ mol^{-1}



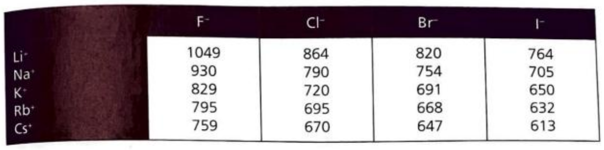
Slide 64: You Try!

- Write an equation to represent the lattice energy of magnesium oxide (ΔHf⦵ MgO= -602 kJ mol^{-1}, ΔH$_a$⦵ Mg(s) = +148 kJ mol^{-1}, second ionization energy of Mg: +1451 kJ mol^{-1})
- Write an equation to represent the second electron affinity of oxygen and comment on the relative values of the first and second values given in the Data Booklet
- Construct a Born-Haber cycle for magnesium oxide.
- Calculate the lattice energy of magnesium oxide.

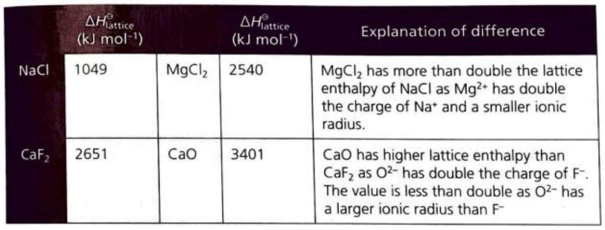
Slide 65: Theoretical Lattice Enthalpies

- Lattice enthalpies can also be calculated from the **ionic model** (assumes only interactions are due to ____________ **forces** between ions).
- Energy needed to ____________ ions depends on
- Ionic charges
- Ionic radius
- Increase in ionic charge increases ____________ between ions
- Increase in ionic radius decreases ____________ between ions

Slide 66: Theoretical Lattice Enthalpies



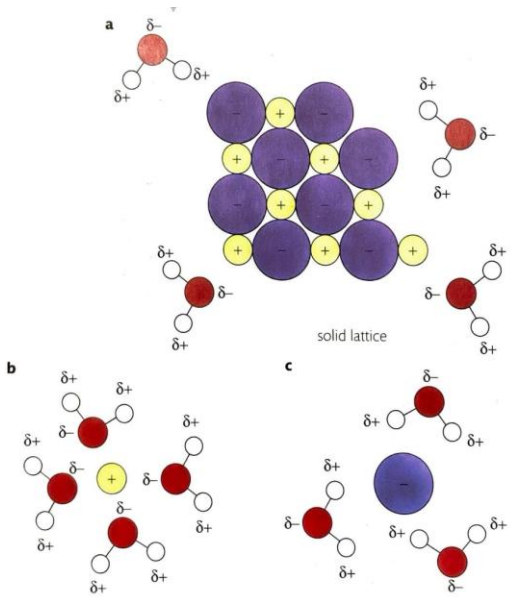
	F⁻	Cl⁻	Br⁻	I⁻
Li⁺	1049	864	820	764
Na⁺	930	790	754	705
K⁺	829	720	691	650
Rb⁺	795	695	668	632
Cs⁺	759	670	647	613



	ΔH⦵lattice (kJ mol⁻¹)		ΔH⦵lattice (kJ mol⁻¹)	Explanation of difference
NaCl	1049	$MgCl_2$	2540	$MgCl_2$ has more than double the lattice enthalpy of NaCl as Mg^{2+} has double the charge of Na^+ and a smaller ionic radius.
CaF_2	2651	CaO	3401	CaO has higher lattice enthalpy than CaF_2 as O^{2-} has double the charge of F^-. The value is less than double as O^{2-} has a larger ionic radius than F^-

Slide 67: Enthalpies of Solution ($\Delta H_{sol}^{\ominus}$)

- **The enthalpy change when one mole of a solute is dissolved in a solvent to infinite dilution under standard ____________.**
- Ions separated from the lattice and ____________ by water molecules are **hydrated**.
- NH_4Cl (s) → NH_4^+(aq) + Cl^-(aq) $\Delta H_{sol}^{\ominus}$ = +14.78 kJ mol^{-1}

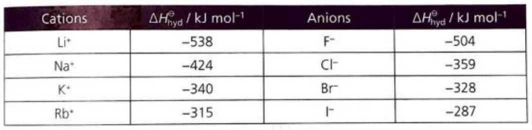


Slide 68: Hydration Enthalpy ($\Delta H_{hyd}^{\ominus}$)

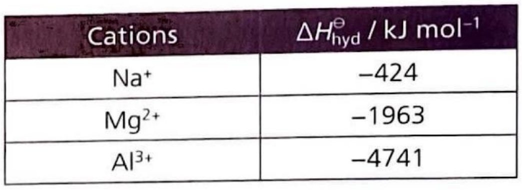
- **The enthalpy change when one mole of gaseous ions is dissolved to form an ____________ dilute solution of one mole of aqueous ions under standard conditions.**
- ____________ of hydration **cannot be measured directly** since both positive and negative ions are present in solution.
- $H^+(g) \rightarrow H^+(aq)$ $\Delta H_{hyd}^{\ominus}$ = -1130 kJ mol^{-1}
- Hydration is an ____________ **process** since there is a force of attraction between ions and polar water molecules.

Slide 69: Hydration Enthalpy

- Values get **less** ____________ as the ionic radius increases down a group
- Hydration enthalpies are ____________ **inversely proportional to the ionic radii.**
- Values get **more** ____________ as the ionic charge increases across period 3



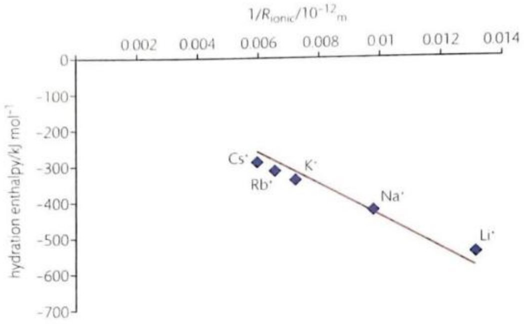
Cations	$\Delta H_{hyd}^{\ominus}$ / kJ mol^{-1}	Anions	$\Delta H_{hyd}^{\ominus}$ / kJ mol^{-1}
Li^+	−538	F^-	−504
Na^+	−424	Cl^-	−359
K^+	−340	Br^-	−328
Rb^+	−315	I^-	−287

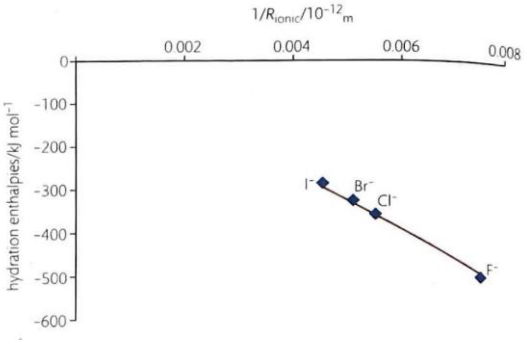


Cations	$\Delta H_{hyd}^{\ominus}$ / kJ mol^{-1}
Na^+	−424
Mg^{2+}	−1963
Al^{3+}	−4741

Slide 70: Hydration Enthalpy

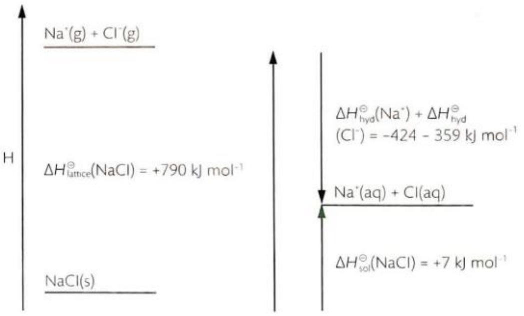
- Hydration _____________ **decrease** as ionic radius **increases**





Slide 71: Relationship between ΔHhyd⊖, ΔHsol⊖, and ΔHlat⊖

- $\Delta H_{sol}^{\ominus}$(NaCl) = $\Delta H_{lat}^{\ominus}$(NaCl) + $\Delta H_{hyd}^{\ominus}$(Na^+) + $\Delta H_{hyd}^{\ominus}$(Cl^-)
- $\Delta H_{sol}^{\ominus}$(NaCl) = +790 - 424 - 359
- $\Delta H_{sol}^{\ominus}$(NaCl) = +7 kJ mol^{-1}
- $\Delta H_{sol}^{\ominus}$(NaCl) reported in the Data Booklet is +3.88 kJ mol^{-1}. This is due to problems calculating small numbers from differences in large numbers.



Slide 72: You Try!

- Use an energy cycle to calculate the enthalpy of solution of potassium chloride ($\Delta H_{hyd}^{\ominus}$(K^+): -340 kJ mol^{-1},$\Delta H_{hyd}^{\ominus}$(Cl^-): -359 kJ mol^{-1})
- Calculate the % inaccuracy of your value by comparing with the theoretical value of $\Delta H_{sol}^{\ominus}$ KCl (+17.22 kJ mol^{-1}) and comment on the disagreement between the two values.