Worksheet 5.1

Chapter 5: Energetics – glossary

**Atomization**  The process in which an element or compound is converted into gaseous atoms.

**Average bond energy**  This is the energy required to break one mole of the same type of bond in the gaseous state averaged over a variety of similar compounds. It is a measure of the strength of the bond.

**Bomb calorimeter**  A device used to measure energy changes (at constant volume) that occur when substances are burnt in excess oxygen in a sealed container.

**Bond enthalpy**  The bond enthalpy is the amount of energy required to break one mole of a particular covalent bond in the gaseous state into gaseous atoms under standard thermodynamic conditions.

**Bond length**  The distance between the nuclei of two atoms covalently bonded together.

**Bond order**  An index of the degree of bonding between two atoms. A single bond has a bond order of 1, a double bond 2 and a triple bond 3. The bonding in benzene and ozone, for example, is an intermediate between a single and double bond so the bond order is 1.5. When a number of resonance structure are needed to represent a molecule the bond order is a weighted average of the bond numbers between the respective atoms in the contributing structures.

**Born–Haber cycle**  The Born–Haber cycle is a special case of Hess’s law for the formation of ionic compounds. It allows the experimental lattice enthalpy to be calculated from other enthalpy changes.

**Calorimetry**  A technique of measuring heat changes in physical processes and chemical reactions.

**Calorimeter**  A piece of apparatus of known heat capacity used for measuring the energy released or absorbed during a chemical reaction from temperature changes in water. When enthalpies of combustion are measured the calorimeter is made from a good conductor, which allows heat from the flame to pass to water. When used to measure the enthalpy changes in solution it is made from an insulator which reduces heat losses from the system.

**Combustion**  A highly exothermic and rapid chemical reaction in which a substance reacts with oxygen during burning. Complete combustion occurs when there is an excess of oxygen and the substance is fully oxidized. Incomplete combustion occurs in a limited supply of air.

**Electron affinity**  The first electron affinity is the enthalpy change when one mole of gaseous atoms accepts a mole of gaseous electrons under standard conditions: \( X (g) + e^- \rightarrow X^- (g) \). Further electron affinities correspond to the acceptance of additional electrons. For oxygen the first electron affinity, \( O (g) + e^- \rightarrow O^- (g) \), is exothermic as the neutral atoms attract the additional electron into its outer energy level. The second electron affinity is endothermic as the extra electron is repelled by the negatively charged ion: \( O^- (g) + e^- \rightarrow O^{2-} (g) \).
Endothermic reactions
An endothermic reaction is one that absorbs heat from the surroundings as a result of forming products with weaker bonds than the reactants. Endothermic reactions have positive $\Delta H$ values.

Enthalpy ($H$)
The amount of heat energy contained in a substance. It is stored in the chemical bonds as potential energy. When substances react, the difference in the enthalpy between the reactants and products (at constant pressure) results in a heat change which can be measured.

Enthalpy change of combustion
The standard enthalpy change of combustion ($\Delta H_C^\theta$) is the enthalpy change for the complete combustion of one mole of a substance in its standard state in excess oxygen under standard conditions.

Enthalpy change of formation
The standard enthalpy change of formation ($\Delta H_f^\theta$) is the enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions.

Enthalpy change of fusion
The enthalpy change ($\Delta H_{\text{fus}}$) that occurs when one mole of a pure solid is converted to a liquid at its melting point.

Enthalpy change of neutralization
The enthalpy change ($\Delta H_{\text{neut}}^\theta$) when one mol of H$^+$ (aq) reacts with 1 mol of OH$^-$ (aq) ions. The reaction is exothermic as bond formation takes place: H$^+$ (aq) + OH$^-$ (aq) $\rightarrow$ H$_2$O (l).

Enthalpy change of solution
The enthalpy change when one mole of a substance is dissolved in a solvent to infinite dilution (in practice, to form a dilute solution) under standard conditions.

Enthalpy level diagram
A diagram that shows the relative enthalpies of the reactants and products in a chemical reaction.

Enthalpy of atomization (of an element)
The standard enthalpy change ($\Delta H_{\text{atom}}^\theta$) when one mole of gaseous atoms is formed from the element in its standard state under standard conditions. For example for carbon: C (graphite) $\rightarrow$ C (g) and for chlorine: $\frac{1}{2}$Cl$_2$ (g) $\rightarrow$ Cl (g).

Enthalpy of vaporization
The enthalpy change ($\Delta H_{\text{vap}}$) that occurs when one mole of a pure liquid is vaporised at its boiling point.

Entropy
Entropy ($S$) is a property which quantifies the degree of disorder or randomness in a system. Ordered states have low $S$, disordered states have high $S$: S (s) < S (l) < S (g).

Entropy change
The change in entropy ($\Delta S$) that accompanies a physical or chemical change is given by the sum of the entropies of the products minus the sum of the entropies of the reactants: $\Delta S^\theta_{\text{reaction}} = \Sigma S^\theta (\text{products}) - \Sigma S^\theta (\text{reactants})$

Exothermic
An exothermic reaction is one that releases heat to the surroundings as a result of forming products with stronger bonds than the reactants. Exothermic reactions have negative $\Delta H$ values.
Gibbs’ free energy change  Gibbs’ free energy ($G$) is the criterion for predicting the spontaneity of a reaction or process: it is related to $\Delta S^\theta_{\text{total}}$. It gives the energy available to do useful work and is related to the enthalpy and entropy changes of the system: 

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\Delta G^\theta_{\text{system}} = \Delta H^\theta_{\text{system}} - T \Delta S^\theta_{\text{system}}.
$$

A negative sign indicates that the reaction is spontaneous under standard conditions.

Heat capacity  The amount of heat energy required to raise the temperature of a substance by one degree kelvin (or Celsius).

Heat energy  The energy transferred between two objects due to a temperature difference between them. The source of heat energy is the kinetic energy of the particles.

Hess’s law  This states that the total enthalpy change for a reaction is independent of the route taken. It is a special case of the law of conservation of energy.

Ionic equation  The simplified equation for a reaction involving ionic substances. Only those ions that actually participate in the reaction are included in the ionic equation; spectator ions are not included.

Ionic model  A model used to calculate theoretical lattice enthalpies which assumes that the only bonding in an ionic solid is the electrostatic forces between the ions. These forces depend on the ionic radii and charges.

Ionization energy (first)  The minimum energy required to remove one mole of electrons from a mole of gaseous atoms to form a mole of univalent cations in the gaseous state. It is the enthalpy change for the process: $X(\text{g}) \rightarrow X^+(\text{g}) + e^-$.  

Joule  The SI unit for energy. Symbol J.

Latent heat  The heat needed to bring about a change in state. The energy is needed to break the bonds between the particles and does not increase their kinetic energy: there is no temperature change.

Lattice  A regular, repeating three-dimensional arrangement of atoms, molecules or ions within a crystal.

Lattice enthalpy  This is the enthalpy change ($\Delta H^\theta_{\text{lat}}$) that occurs when one mole of a solid ionic compound is separated into gaseous ions under standard conditions. For example, for alkali metal halides: $\text{MX (s)} \rightarrow \text{M}^+(\text{g}) + \text{X}^- (\text{g}) \Delta H^\theta_{\text{lat}}$.

Limiting reactant  The reactant that is completely consumed, or used up, when a reaction goes to completion.

Random errors  Random errors in thermochemistry are caused by uncertainties in the readability of the measuring instruments and the effects of changes in the surroundings, such as temperature variations.

Specific heat capacity  The specific heat capacity is the amount of heat energy required to raise the temperature of unit mass (e.g. 1 kg or 1 g) of a substance, by 1°C or 1 K.
**Spontaneous reaction**  A reaction which will occur when the reactants are mixed together under standard conditions. It does not require an input of heat energy and is accompanied by a decrease in Gibbs’ free energy.

**Stability**  A compound is described as being stable if it does not tend to decompose into its elements or into other compounds.

**Standard conditions**  Standard conditions are:
- pressure = 101.3 kPa, temperature = 298 K and concentration = 1 mol dm\(^{-3}\)

**Standard enthalpy change of reaction**  The standard enthalpy change (\(\Delta H^0\)) is the heat energy transferred under standard conditions (pressure 101.3 kPa, temperature 298 K). Only \(\Delta H^0\) can be measured, not \(H\) for the initial or final state of a system.

**Standard state**  The standard state of an element or compound is its most stable state under the specified conditions. For example, the standard state of H\(_2\)O at 298 K and 1 atm. is H\(_2\)O (l).

**States of matter**  Solid, liquid and gas are the three states of matter in which all substances can exist, depending on the temperature and pressure.

**System**  A term used to describe the material or mixture of chemicals being studied. Everything outside the system is the surroundings. An open system exchanges matter and energy with its surroundings. A closed system exchanges energy but not matter. An isolated system exchanges neither matter nor energy.

**Systematic errors**  These occur when there is an error in the experimental procedure, such as heat losses in an exothermic reaction which lead to smaller temperature changes.

**Thermochemical equation**  A chemical equation that describes the reaction occurring and gives the associated enthalpy change. State symbols must be shown as \(\Delta H^0\) depends on the state of the reactants or products.

**Thermochemistry**  The study of heat changes occurring during chemical reactions.

**Thermodynamics**  The laws which govern energy transfers and the direction of chemical and physical changes. The first law of thermodynamics is the conservation of energy. The second law states that changes occur in the universe when the total entropy increases.