Access to Higher Education Unit

This unit forms part of an Access to HE Diploma. If delivering the graded version of this unit, please refer to the Provider Handbook for details on grading descriptors and the application of these across units within your programme.

Unit Title: Fundamental Principles of Physical Chemistry

Graded Unit Reference Number: GA33CHE

Ungraded Unit Reference Number: UA33CHE

Module: Chemistry

Level: Three (3)

Credit Value: Six (6)

Minimum Guided Learning Hours: 60

Units barred for selection against this unit:

- **Energetics (GA33CHE12 / UA33CHE12)**
- **Chemical and Acid-base Equilibria (GA33CHE11 / UA33CHE11)**
- **Kinetics and Redox Systems (GA33CHE13 / UA33CHE13)**
- **Thermodynamics (GA36CHE19 / UA36CHE19)**

Learning Outcome (The Learner will):	Assessment Criterion (The Learner can):
1. Understand the factors that influence the rate of chemical reactions	1.1 Use collision theory to explain factors that affect the rate of a chemical reaction.
	1.2 Analyse Maxwell–Boltzmann distribution curves to explain the effect of temperature on reaction rate.
	1.3 Investigate the effect of changing reaction conditions on the rate of a chemical reaction.
2. Understand enthalpy changes in exothermic and endothermic reaction	2.1 Interpret enthalpy profile diagrams for exothermic and endothermic reactions, including activation energy.
	2.2 Describe standard conditions and standard enthalpy changes of combustion, formation, and reaction.
	2.3 Calculate enthalpy changes of reaction using average bond enthalpy values or experimental data.

		2.4	Carry out an experiment to determine the enthalpy change of a reaction.
3.	Apply Le Chatelier's principle to systems in dynamic equilibrium	3.1	State Le Chatelier's principle and apply it to explain the effects of changes in conditions on the position of equilibrium.
		3.2	Investigate the effect of changing conditions on a chemical system at equilibrium.
4.	Understand the theory and reactions of acids and bases	4.1	Define acids and bases using the Brønsted–Lowry theory.
		4.2	Explain how the pH scale is used to measure acidity and alkalinity.
		4.3	Explain the reactions of acids with alkalis, metals, and carbonates.
		4.4	Carry out acid–base reactions and use pH measurements to analyse acidity and alkalinity.
5.	Understand simple redox processes	5.1	Explain oxidation and reduction in terms of electron transfer and changes in oxidation state.
		5.2	Assign oxidation states to atoms in compounds and ions.
		5.3	State Hess' Law and use enthalpy cycles to determine enthalpy changes indirectly.
		5.4	Carry out a redox reaction and identify oxidation and reduction processes

Indicative Content Learners could cover the following: (The information below is provided for guidance only and is not mandatory)	
LO1	AC1.1 **Collision theory principles:** particles must collide with sufficient energy (activation energy) and correct orientation for reaction to occur. **Factors affecting rate:** temperature (increases particle kinetic energy and collision frequency), concentration/pressure (increases number of particles per unit volume), surface area (greater exposure for collisions), catalysts (lower activation energy, alternative pathway). **Role of activation energy:** minimum energy required for successful collisions; effect on reaction speed. AC1.2 **Distribution of kinetic energy:** spread of molecular energies at a given temperature; most probable energy.

	Effect of temperature change: increased temperature shifts curve to higher energies, increasing proportion of particles above activation energy. **Interpretation of graphs:** area under curve represents number of particles; comparison of curves at different temperatures. AC1.3 **Experimental methods:** measuring rate via gas production, mass loss, colour change, precipitation (e.g., disappearing cross experiment). **Variables:** independent, dependent, control variables. **Data collection and analysis:** recording measurements, plotting graphs (e.g., concentration vs time), calculating rate (initial rate or mean rate). **Evaluation:** reliability, accuracy, sources of experimental error, improvements and safety considerations.
LO2	AC2.1 **Exothermic and endothermic reactions:** energy release vs energy absorption; relative energy levels of reactants and products. **Activation energy:** energy barrier shown on profile diagrams. **Energy changes (ΔH):** negative ΔH for exothermic, positive ΔH for endothermic reactions. AC2.2 **Standard conditions:** 298 K (25°C), 100 kPa pressure, 1 mol dm^{-3} concentration. **Standard enthalpy changes:** combustion, formation, and reaction definitions and examples. **Notation:** use of ΔH° and appropriate units (kJ mol^{-1}). AC2.3 **Bond enthalpy calculations:** breaking bonds (endothermic) vs forming bonds (exothermic). **Use of experimental data:** q = mcΔT calculations for calorimetry. **Units and conversions:** energy, mass, temperature change, moles. AC2.4 **Calorimetry experiments:** simple calorimeter setup, measuring temperature change. **Procedure considerations:** insulation, accurate measurement, minimising heat loss. **Data handling:** calculating ΔH from experimental results. **Evaluation:** errors (heat loss, incomplete reaction), reliability, suggested improvements.

LO3	AC3.1 **Definition:** systems at equilibrium respond to counteract changes imposed on them. **Effect of changing conditions:** • Temperature (favouring endothermic/exothermic direction) • Pressure (gases: shift to fewer/more moles) • Concentration (shift to consume added/reactant/product) **Dynamic equilibrium:** forward and reverse reactions occurring at equal rates in a closed system. AC3.2 **Practical examples:** colour changes in equilibrium mixtures (e.g., cobalt chloride equilibrium), chromate/dichromate system. **Manipulation of conditions:** changing temperature, concentration, or pressure. **Observation and interpretation:** identifying shifts in equilibrium position based on experimental evidence. **Evaluation:** linking observations to theory, considering limitations and accuracy.
LO4	AC4.1 **Definitions:** acids as proton donors, bases as proton acceptors. **Conjugate acid–base pairs:** formation and identification in reactions. AC4.2 **Definition of pH:** measure of hydrogen ion concentration. **Scale interpretation:** strong vs weak acids/bases; logarithmic nature of scale. **Measurement methods:** pH meters, universal indicator, pH paper. AC4.3 **Acid reactions:** • With metals (producing hydrogen gas) • With bases/alkalis (neutralisation forming salt and water) • With carbonates (producing carbon dioxide) **Word and symbol equations** for typical reactions. AC4.4 **Titration techniques:** use of burettes, pipettes, indicators. **pH measurements:** monitoring acidity/alkalinity. **Data analysis:** determining equivalence point, calculating concentration. **Safety and evaluation:** handling acids/alkalis, identifying errors and improvements.

LO5	AC5.1 **Electron transfer:** oxidation as loss of electrons, reduction as gain. **Oxidation states:** increase (oxidation) and decrease (reduction). AC5.2 **Rules:** common oxidation states (e.g., group 1 = +1, oxygen = –2). **Application:** determining oxidation states in compounds and ions. AC5.3 **Hess' Law principle:** total enthalpy change independent of pathway. **Enthalpy cycles:** construction and use to determine enthalpy changes indirectly. **Calculation methods:** using formation or combustion data. AC5.4 **Examples:** displacement reactions, redox titrations. **Identification:** recognising oxidation and reduction processes within reactions. **Observations:** colour changes, gas evolution, metal displacement. **Evaluation:** linking experimental outcomes to redox theory, assessing accuracy and reliability.