Grade 12

A: Reaction Kinetics

(Introduction)

It is expected that students will:

A1. recognize that reactions proceed at different rates

A2. describe or calculate rate in terms of some quantity (produced or consumed) per unit of time

A3. determine rate of a reaction from experimental data, and describe how rate changes over time as the reaction proceeds

A4. identify properties that could be monitored in order to determine a reaction rate

A5. recognize some of the factors that control reaction rates

A6. compare and contrast factors affecting the rates of both homogeneous and heterogeneous reactions

A7. discuss situations in which the rate of reaction must be controlled

B: Reaction Kinetics

(Collision Theory)

It is expected that students will:

B1. demonstrate an awareness of the following:

- reactions are the result of collisions between reactant particles
- not all collisions are successful
- sufficient kinetic energy (KE) and favourable geometry are required to increase the rate of a reaction one must increase the frequency of successful collisions
- a KE distribution curve can explain how changing temperature or adding a catalyst changes the rate
- energy changes are involved in reactions as bonds are broken and formed

B2. describe the activated complex in terms of its potential energy (PE), stability, and structure

B3. define activation energy

B4. describe the relationship between activation energy and rate of reaction

B5. describe the changes in KE and PE as reactant molecules approach each other

B6. draw and label PE diagrams for both exothermic and endothermic reactions, including:
- the relative PE of reactants, products, and activated complex
- activation energy
- ΔH

B7. relate the sign of ΔH to whether the reaction is exothermic or endothermic

B8. write chemical equations that describe energy effects in 2 ways:
- a chemical equation that includes the energy term (thermochemical equation)
- a chemical equation using ΔH notation

B9. use collision theory to explain the effect of the following factors on reaction rate
- nature of reactants
- concentration
- temperature
- surface area

C: Reaction Kinetics

(Reaction Mechanisms and Catalysts)

It is expected that students will:

C1. explain why most reactions involve more than one step

C2. describe a reaction mechanism as the series of steps (collisions) that result in the overall reaction and describe the role of the rate determining step

C3. define catalyst

C4. compare and contrast the PE diagrams for a catalyzed and uncatalyzed reaction in terms of:
- reaction mechanism
- ΔH
- activation energy

C5. identify reactant, product, reaction intermediate, activated complex, and catalyst from a given reaction mechanism

C6. describe the uses of specific catalysts in a variety of situations including Pt in catalytic converters in automobiles

D: Dynamic Equilibrium

(Introduction)

It is expected that students will:

D1. describe the reversible nature of most chemical reactions

D2. identify the reversible pathways of a chemical reaction on the PE diagram

D3. relate the changes in rates of the forward and reverse reactions to the changing concentrations of the reactants and products as equilibrium is established

D4. describe chemical equilibrium as a closed system at constant temperature:
- whose macroscopic properties are constant
- where the forward and reverse reaction rates are equal
- that can be achieved from either direction
- where the concentrations of reactants and products are constant

D5. describe the dynamic nature of chemical equilibrium

D6. state that a system not at equilibrium will tend to move toward a position of equilibrium

D7. determine entropy and enthalpy changes from a chemical equation (qualitatively)

D8. state that systems tend toward a position of minimum enthalpy and maximum entropy (randomness)

D9. predict the result when enthalpy and entropy factors:
- both favour the products
- both favour the reactants
- oppose one another

E: Dynamic Equilibrium

(Le Châtelier's Principle)

It is expected that students will:

E1. describe the term shift as it applies to equilibria

E2. apply Le Châtelier's principle to the shifting of equilibrium involving the following:

- temperature change
- concentration change
- volume change of gaseous systems
- adding an inert gas to a gaseous system

E3. explain the above shifts using the concepts of reaction kinetics

E4. identify the effect of a catalyst on dynamic equilibrium

E5. apply the concept of equilibrium to a commercial or industrial process

F: Dynamic Equilibrium

(The Equilibrium Constant)

It is expected that students will:

F1. gather and interpret data on the concentration of reactants and products of a system at equilibrium

F2. Write the expression for the equilibrium constant when given the equation for either a homogeneous or heterogeneous equilibrium system. Explain why certain terms are not included in the expression.

F3. relate the equilibrium position to the value of Keq and vice versa

F4. predict the effect (or lack of effect) on the value of Keq of changes in the following factors: temperature, pressure, concentration, surface area, and catalyst

F5. perform calculations involving the value of Keq and the equilibrium concentration of all species

F6. perform calculations involving the value of Keq, the initial concentrations of all species and one equilibrium concentration

F7. calculate the equilibrium concentrations or amounts of all species given the value of Keq and the initial concentrations

F8. determine whether a system is at equilibrium, and if not, in which direction it will shift to reach equilibrium when given a set of concentrations for reactants and products

G: Solubility Equilibria

(Concept of Solubility)

It is expected that students will:

G1. classify a solution as ionic or molecular given its conductivity or the formula of its solute

G2. describe the conditions necessary to form a saturated solution

G3. describe solubility as the concentration of a substance in a saturated solution

G4. use appropriate units to represent the solubility of substances in solutions

G5. determine the solubility of a compound in solution

G6. describe the equilibrium that exists in a saturated solution

G7. write a net ionic equation that describes the equilibrium in a saturated solution

G8. calculate the concentration of the positive and negative ions given the concentration of a solute dissolved in a solution

H: Solubility Equilibria

(Solubility and Precipitation)

It is expected that students will:

H1. describe a compound as having high or low solubility relative to 0.1 M by using a solubility chart

H2. use a solubility chart to predict if a precipitate will form when two solutions are mixed, and identify the precipitate

H3. write a formula equation, complete ionic equation, and net ionic equation that represent a precipitation reaction

H4. use a solubility chart to predict if ions can be individually separated from solution through precipitation reactions, and outline an experimental procedure which includes

- compound added
- precipitate formed
- method of separation

H5. predict qualitative changes in the solubility equilibrium upon the addition of a common ion, or the removal of an ion

H6. identify an unknown ion through experimentation involving a qualitative analysis scheme

H7. devise a procedure by which the calcium and or magnesium ions can be removed from hard water

I: Solubility Equilibria

(Quantitative Aspects)

It is expected that students will:

I1. describe the Ksp expression as a specialized Keq expression

I2. write a Ksp expression for a solubility equilibrium

I3. calculate the Ksp for a compound when given its solubility.
For example: AgCl, Ag_2S, $PbCl_2$

I4. calculate the solubility of a compound from its Ksp

I5. predict the formation of a precipitate by comparing the trial Ksp to the accepted Kspvalue using specific data

I6. calculate the maximum allowable concentration of one ion given the Ksp and the concentration of the other ion just before precipitation occurs

I7. demonstrate and describe a method for determining the concentration of a specific ion

J: Acids, Bases, and Salts

(Properties and Definitions)

It is expected that students will:

J1. identify acids and bases through experimentation

J2. list general properties of acids and bases

J3. write balanced equations representing the neutralization of acids by bases in solution

J4. define Arrhenius acids and bases

J5. write names and formulae of some common household acids and bases

J6. define Brönsted-Lowry acids and bases

J7. identify and label Brönsted-Lowry acids and bases in an equation

J8. write balanced equations representing the reaction of acids or bases with water

J9. identify an H_30^+ ion as a protonated H_20 molecule that can be represented in shortened form as $H^+(aq)$

J10. define conjugate acid-base pair

J11. identify the conjugate of a given acid or base

J12. show that in any Brönsted-Lowry acid-base equation there are two conjugate pairs present

K: Acids, Bases, and Salts

(Strong and Weak Acids and Bases)

It is expected that students will:

K1. relate electrical conductivity in a solution to the total concentration of ions in the solution

K2. classify an acid or base in solution as either weak or strong by comparing conductivity

K3. define a strong acid and a strong base

K4. define a weak acid and a weak base

K5. write equations to show what happens when strong and weak acids and bases are dissolved in water

K6. compare the relative strengths of acids or bases by using a table of relative acid strengths

K7. identify and explain why the strongest acid in aqueous solutions is H_30^+ and the strongest base in aqueous solutions is OH^-

K8. predict whether products or reactants are favoured in an acid-base equilibrium by comparing the strength of the two acids (or two bases)

K9. compare the relative concentrations of H_30^+ (or OH^-) between two acids (or between two bases) using their relative positions on an acid strength table

K10. define amphiprotic

K11. identify chemical species that are amphiprotic

K12. describe situations in which H_20 would act as an acid or base

L: Acids, Bases, and Salts

(Kw, pH, pOH)

It is expected that students will:

L1. write equations representing the ionization of water using either H_30^+ and OH^- or H^+ and OH^-

L2. write the equilibrium expression for the ion product constant of water, (water ionization constant), Kw

L3. predict the effect of the addition of an acid or base to the equilibrium system: $2H_2O(\ell) \leftrightarrow H_30^+(aq)+OH^-(aq)$

L4. state the relative concentrations of H_30^+ and OH^- in acid, base, and neutral solutions

L5. state the value of Kw at 25°C

L6. describe and explain the variation of the value of Kw with temperature

L7. calculate the concentration of $[H_30^+]$(or $[OH^-]$) given the other, using Kw

L8. describe the pH scale with reference to everyday solutions

L9. define pH and pOH

L10. define pKw , give its value at 25°C, and its relation to pH and pOH

L11. calculate pH, and pOH from $[H_30^+]$ or $[OH^-]$

L12. calculate $[H_30^+]$ and $[OH^-]$ from pH or pOH

M: Acids, Bases, and Salts

(Ka and Kb Problem Solving)

It is expected that students will:

M1. write Ka and Kb equilibrium expressions for weak acids or bases

M2. relate the magnitude of Ka (the acid ionization constant) or Kb (the base ionization constant) to the strength of the acid or base

M3. given the Ka, Kb, and initial concentration, calculate any of the following:

- $[H_30^+]$
- $[OH^-]$
- pH
- pOH

M4. calculate the value of Kb for a base given the value of Ka for its conjugate acid (or vice versa)

M5. calculate the value of Ka or Kb given the pH and initial concentration

M6. calculate the initial concentration of an acid or base given the appropriate Ka, Kb, pH or pOH values

N: Acids, Bases, and Salts

(Hydrolysis of Salts)

It is expected that students will:

N1. write a dissociation equation for a salt in water

N2. write net ionic equations representing the hydrolysis of salts

N3. predict whether a salt solution would be acidic, basic, or neutral (compare Ka and Kb values where necessary)

N4. determine whether an amphiprotic ion will act as a base or an acid in solution (compare Ka and Kb values where necessary)

N5. Calculate the pH of a salt solution from relevant data, assuming that the predominant hydrolysis reaction is the only reaction determining the pH

O: Acids, Bases, and Salts

(Indicators)

It is expected that students will:

O1. describe an indicator as a mixture of a weak acid and its conjugate base, each with distinguishing colours

O2. describe the term transition point of an indicator, including the conditions that exist in the equilibrium system

O3. describe the shift in equilibrium and resulting colour changes as an acid or a base is added to an indicator

O4. predict the approximate pH at the transition point using the Ka value of an indicator

O5. predict the approximate Ka value for an indicator given the approximate pH range of the colour change

O6. match an indicator colour in a solution with an approximate pH, using a table of indicators

P: Acids, Bases, and Salts

(Neutralizations of Acids and Bases)

It is expected that students will:

P1. demonstrate an ability to design and perform a titration experiment involving the following:

- primary standards
- standardized solutions
- titration curves
- indicators selected so the transition point of the indicator coincides with the equivalence point of the titration reaction

P2. calculate the concentration of an acid or base using titration data or similar data such as grams or moles

P3. calculate the volume of an acid or base of known molarity needed to completely react with a given amount of base or acid

P4. write formula, complete ionic, and net ionic equations for:

- a strong acid reacting with a strong base (neutralization)

- a weak acid reacting with a strong base
- a strong acid reacting with a weak base

P5. calculate the pH of a solution formed when a strong acid is mixed with a strong base

P6. explain the difference between the equivalence point (stoichiometric point) of a strong acid-strong base titration and the equivalence point of a titration involving a weak acid-strong base or strong acid-weak base

Q: Acids, Bases, and Salts

(Buffer Solutions)

It is expected that students will:

Q1. describe the function of a buffer solution to resist changes in pH. It must be able to buffer the addition of small amounts of strong acid and it must also buffer the addition of small amounts of strong base

Q2. describe the composition of an acidic buffer and a basic buffer

Q3. outline a procedure to prepare a buffer solution

Q4. identify the limitations in buffering action (as being the amount of either the weak acid (or base) or its conjugate base (or acid)

Q5. describe qualitatively how the buffer equilibrium shifts as small quantities of acid or base are added to the buffer; the stress being the change in the concentration of the stronger acid (H_3O^+) or base (OH^-)

Q6. describe common buffer systems present in industrial, environmental, or biological systems including the common blood buffer

R: Acids, Bases, and Salts

(Acid Rain)

It is expected that students will:

R1. predict whether a metal or nonmetal oxide will produce acidic or basic solutions; write equations representing the formation of the acidic or basic solutions

R2. describe the pH conditions (5.0 and lower) required for rain to be called acid rain

R3. relate the pH (about 5.6) of normal rain to the presence of dissolved CO_2

R4. describe common sources of NOx (automobile engines) and SOx (smelters of sulphide ores and fuels containing sulphur)

R5. discuss general environmental problems associated with acid rain

S: Oxidation-Reduction

(Introduction)

It is expected that students will:

S1. define and identify the following:
- oxidation
- reduction
- oxidizing agent
- reducing agent
- half-reaction
- redox reaction

S2. determine the following:
- the oxidation number of an atom in a chemical species
- the change in oxidation number an atom undergoes when it is oxidized or reduced
- whether an atom has been oxidized or reduced by its change in oxidation number

S3. relate change in oxidation number to gain or loss of electrons

S4. from data for a series of redox reactions, create a table of reduction half-reactions

S5. identify the relative strengths of oxidizing and reducing agents from their positions on a half-reaction table

S6. use a table of reduction half-reactions to predict whether a spontaneous redox reaction will occur between any two species

T: Oxidation-Reduction

(Balancing Redox Equations)

It is expected that students will:

T1. balance a half-reaction in solution (acid, base, neutral)

T2. balance a net ionic redox equation in acid and base solution

T3. write the equations for reduction and oxidation half-reactions given a redox reaction

T4. identify reactants and products for several redox reactions performed in a laboratory and write balanced equations

T5. select a suitable reagent to be used in a redox titration in order to determine the concentration of a species, and be familiar with a few common reagents used in redox titrations, including acidified permanganate and dichromate

T6. determine the concentration or amount of a species from redox titration data, or similar data such as grams or moles

U: Oxidation-Reduction

(Electrochemical Cells)

It is expected that students will:

U1. define, construct, and label the parts of an electrochemical cell

U2. determine the half-reactions that take place at each electrode of an electrochemical cell, and use these reactions to predict typical observations as well as the overall reaction

U3. predict the direction of movement of each type of ion in the cell

U4. predict the direction of flow of electrons in an external circuit

U5. predict what happens to the mass of each electrode as the cell operates

U6. predict the voltage (potential) of the cell when equilibrium is reached

U7. determine voltages of half-reactions by analyzing the voltages of several cells. Recognize that reduction half cell voltages are relative to the standard hydrogen half cell (voltage = 0V)

U8. recognize that E° of an electrochemical cell represents the cell potential at standard conditions

U9. predict the voltage (E°) of an electrochemical cell using the table of standard reduction half-cells

U10. predict the spontaneity of the forward or reverse reaction from the E° of a redox reaction

U11. demonstrate a general understanding of applications of electrochemical cells, including lead-acid storage batteries, alkali cells and hydrogen-oxygen fuel cells

V: Oxidation-Reduction

(Corrosion)

It is expected that students will:

V1. describe the conditions necessary for corrosion of metals (spontaneous oxidation) to occur

V2. describe the process of metal corrosion in electrochemical terms

V3. suggest several methods of preventing or inhibiting corrosion of a metal, and explain why these methods work.

V4. describe and explain the principle of cathodic protection

W: Oxidation-Reduction

(Electrolytic Cells)

It is expected that students will:

W1. define electrolysis and electrolytic cell

W2. design and label the parts of an electrolytic cell used for the electrolysis of a molten binary salt such as NaCl liquid

W3. design and label the parts of an electrolytic cell capable of electrolyzing an aqueous salt such as KI aqueous (use of overpotential effect not required)

W4. predict the direction of flow of all ions in the cell and electrons in the external circuit

W5. write the half-reactions occurring at each electrode and predict observations based on these half reactions

W6. write the overall cell reaction and predict the minimum voltage required for it to operate under standard conditions

W7. demonstrate an understanding of the principles involved in simple electroplating

W8. construct an electrolytic cell capable of electroplating an object

W9. demonstrate a general understanding of applications of electrolytic cells in metal

refining processes including refining of Zn and Al