COURSE NUMBER:	CH1120
COURSE TITLE:	Chemistry

COURSE DESCRIPTION:

This in an introductory course designed to give students knowledge and understanding of the fundamental chemical concepts which will form the basis for further studies in science and technology.

PREREQUISITES:	None
CO-REQUISITES:	None
CREDIT VALUE:	Four (4)
COURSE HOURS PER WEEK:	Three (3)
LAB HOURS PER WEEK:	Two (2) (NL only) Three (3) (Qatar only)
SUGGESTED TEXT:	To be determined by instructor
LEARNING RESOURCES:	To be determined by instructor

MAJOR TOPICS:

1.0 Atomic Structure
2.0 Periodic Table
3.0 Chemical Bonding and Nomenclature
4.0 Mole Calculations
5.0 Chemical Reactions
6.0 Kinetic Molecular Theory and Gas Laws

LEARNING OBJECTIVES:

The expected learning outcome is that the student will be able to:

1.0 Atomic Structure

1.1 Fundamentals of Measurement
- 1.1.1 Define the Canadian System of Measurements (SI)
- 1.1.2 Define accuracy, precision and significant figures
- 1.1.3 Perform calculations involving significant figures

1.2 Composition of Matter
- 1.2.1 Elements and Compounds
  - 1.2.1.1 Distinguish between elements and compounds
  - 1.2.1.2 Describe how a chemical symbol is written
  - 1.2.1.3 State the chemical symbol given the name, and name given the symbol, for the following elements of the Period Table: H, He, Li, Be, B, C, N, O, F, Ne, Na, Mg, Al, Si, P, S, Cl, Ar, K, Ca, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ge, Ag, Br, Pd, Cd, I, Cs, Ba, W, Pt, Hg, Pb, Rn, Sn, Au, U

- 1.2.2 Law of Constant Composition
  - 1.2.2.1 State the Law of Constant Composition

- 1.2.3 Law of Multiple Proportion
  - 1.2.3.1 State the Law of Multiple Proportion

- 1.2.4 Define atom

1.3 Subatomic Particles
- 1.3.1 State that an atom is composed of protons, electrons and neutrons
- 1.3.2 State that the protons are positively charged with a relative mass = 1
- 1.3.3 State that the neutrons are neutral, with a relative mass = 1
- 1.3.4 State that the electrons are negatively charged with negligible mass
- 1.3.5 State that an atom is composed of an extremely small central nucleus containing protons and neutrons surrounded by electrons
- 1.3.6 State that an atom has an equal number of protons and electrons and is therefore electrically neutral

1.4 Atomic Number and Mass
- 1.4.1 Define Atomic Number as the number of protons in the nucleus
- 1.4.2 Define Mass Number as the total number of protons and neutrons in the nucleus
- 1.4.3 Use the Periodic Table to determine the atomic number of an element
- 1.4.4 Use any suitable combination of Atomic Number, Mass Number, number of protons, electrons and neutrons to deduce the others
- 1.4.5 Define average atomic mass
- 1.4.6 Calculate average atomic mass from given values of percent abundance

1.5 Isotopes
- 1.5.1 Use information in 1.4 to explain and illustrate the occurrence of isotopes
- 1.5.2 Explain what makes a nucleus stable
- 1.5.3 Give examples of uses of isotopes

1.6 Quantum Mechanical Model of the atom
- 1.6.1 State that electrons can be found in energy levels and sub-levels

1.6.2 Define s, p, d and f suborbitals
1.6.3 Apply the Aufbau Principle to the concept of orbitals
1.6.4 State that in the Quantum Mechanical Model of the atom, electrons can be grouped according to their ease of removal from the atom, which depends on their distance from the nucleus
1.6.5 Write the simple electronic configuration of an atom in terms of electrons based in sub-levels, given the atomic number concerned
1.6.6 Define valence electrons
1.6.7 Predict the number of valence electrons in an atom
1.6.8 Define ion
1.6.9 List two types of ions
1.6.10 Name positive ions as cations
1.6.11 Name negative ions as anions

2.0 Periodic Table

2.1 Structure of the Periodic Table

2.1.1 State that elements are arranged in the Periodic Table in order of increasing atomic number
2.1.2 State that horizontal rows of elements are called periods and vertical columns are called groups or families, and that elements in the same group have similar properties
2.1.3 State that Groups IA through VIIIA are called main groups; that elements in Groups IB through VIIIB are called transition metals; and that the two rows of elements placed below the main part of the table are called lanthanons and actinons
2.1.4 State that Group IA are known as the alkali metals, Group IIA are known as the alkaline earth metals, Group VIIA are known as the halogens, and that Group VIIIA are known as the noble gases
2.1.5 Classify elements as main group or transition elements
2.1.6 Classify main group elements to a particular group or family

2.2 Metals and Nonmetals

2.2.1 State that all elements can be classified as either metals, non-metals, or semimetals (metalloids)
2.2.2 State the properties of metals, non-metals and semimetals
2.2.3 Classify elements as metals, nonmetals or semimetals based on their position in the periodic table, or on given properties
2.2.4 Relate the state of the elements to the degree of metallic character
2.2.5 Define semiconductor
2.2.6 State which semimetals are semiconductors

2.3 Families of Elements

2.3.1 State basic physical and chemical properties of the alkali metals, alkaline earth metals, halogens and noble gases

2.4 Periodic Variation

2.4.1 Define density
2.4.2 Describe periodic trends in density
2.4.3 Relate simple electronic configuration to position in the Periodic Table
2.4.4 Identify stable (noble gas) electronic configurations
2.4.5 Describe how metals lose valence electrons and nonmetals gain valence electrons to obtain ions with a stable electron configuration
2.4.6 Define atomic radius
2.4.7 Define ionic radius
2.4.8 Define ionization energy
2.4.9 Define electron affinity
2.4.10 Describe and explain the trends in valence electrons, metallic/nonmetallic character, ionic radius, atomic radius, ionization energy and electron affinity across a period and within a group of the Periodic Table
2.4.11 Predict physical properties of elements based on trends within a group

3.0 Chemical Bonding and Nomenclature

3.1 States of Matter

3.1.1 Define solids, liquids, and gases
3.1.2 Define phase change
3.1.3 Define melting, freezing, and melting and freezing points
3.1.4 Define vapour and vapour pressure
3.1.5 Define boiling and boiling point
3.1.6 Explain evaporation, condensation, sublimation, deposition, melting, freezing, and boiling in terms of Kinetic Molecular Theory

3.2 Octet Rule

3.2.1 State that atoms bond in such a way as to obtain a stable electron configuration
3.2.2 Define the Octet Rule

3.3 Ionic Bonding

3.3.1 Ions - Charge on Main Group Elements and Common Transition Elements
3.3.1.1 Describe how atoms form ions through electron transfer and how ions are held together in the crystal lattice

3.3.2 Definition of the Ionic Bond
3.3.2.1 State that ionic bonds are formed by the transfer of electrons from a metallic element to a non-metallic element
3.3.2.2 State that compounds formed between metals and nonmetals are generally ionic compounds
3.3.2.3 State that the smallest unit of an ionic compound is the formula unit

3.3.3 Ionic Trends

3.3.3.1 Describe and explain the trends in melting point, boiling point and conductivity of ionic compounds

3.3.4 Electron-Dot Formula of Ionic Compounds
- 3.3.4.1 Draw electron-dot formulae for atoms
- 3.3.4.2 Draw electron-dot formulae for binary ionic compounds

3.3.5 Chemical Formulae and Nomenclature of Ionic Compounds
- 3.3.5.1 Write formulae for binary ionic compounds given the systematic name, and give the names knowing the formulae
- 3.3.5.2 Name common polyatomic ions
- 3.3.5.3 Write formulae for compounds containing polyatomic ions, using appropriate prefixes and suffixes
- 3.3.5.4 Define hydrate
- 3.3.5.5 Write formulae and names for hydrates

3.4 Covalent Bonding

3.4.1 Definition of the Covalent Bond
- 3.4.1.1 Describe the covalent bond in terms of mutual attraction of nuclei for shared electrons
- 3.4.1.2 State that covalent bonds form between nonmetallic elements, and that compounds formed between non-metallic elements are known as covalent compounds
- 3.4.1.3 State that the smallest unit of a covalent compound is a molecule

3.4.2 Electron-Dot Formulae of Covalent Compounds
- 3.4.2.1 Draw electron-dot formulae for covalent compounds

3.4.3 Chemical Formulae and Nomenclature of Covalent Compounds
- 3.4.3.1 Write formulae for binary covalent compounds given the systematic name, and give names knowing the formulae

3.4.4 Multiple Bonds
- 3.4.4.1 Identify covalent bonds as single (one shared pair of electrons), double (two shared pairs of electrons), or triple (three shared pairs of electrons)
- 3.4.4.2 Write electron-dot structures for covalent compounds with multiple bonds

3.4.5 Covalent Trends
- 3.4.5.1 Describe and explain trends in melting points, boiling points and conductivity of covalent compounds

3.4.6 Polar Covalent Bonding
- 3.4.6.1 Define electronegativity and describe the trends in electronegativity within the Periodic Table

3.4.6.2 Depending on the electronegativity, explain why a bond is purely ionic, purely covalent or polar covalent
3.4.6.3 Identify the atoms which are partially negative and which are partially positive in a binary covalent bond

3.5 Polarity of Molecules
3.5.1 State molecular polarity may be determined by (a) the presence of lone pairs of electrons on the central atom in a molecule, or (b) the presence of different elements surrounding the central atom, or (c) binary molecules involving two different elements
3.5.2 Sketch molecular shapes for binary covalent compounds using VSEPR theory

3.6 Intermolecular Forces
3.6.1 Distinguish between intermolecular and intramolecular forces
3.6.2 Describe:
3.6.2.1 Dispersion forces
3.6.2.2 Dipole-dipole forces
3.6.2.3 Hydrogen bonding forces

3.6.3 Determine the type of intermolecular forces that exist between molecules
3.6.4 Explain trends in physical properties, such as melting and boiling points, by considering the type of intermolecular forces between molecules
3.6.5 Explain, using hydrogen bonding, why water has unique properties such as its density as a solid, high melting and boiling points and thermal properties

3.7 Acids and Bases
3.7.1 Properties of Acids and Bases
3.7.1.1 List properties of acids and bases
3.7.1.2 Explain the properties of acids in terms of hydrogen ions
3.7.1.3 Explain the properties of bases in terms of hydroxide ions

3.7.2 Nomenclature of Acids and Bases
3.7.2.1 Write formulas and names of common acids and bases

4.0 Mole Calculations

4.1 Avogadro's Number - The Mole
4.1.1 Define atomic mass relative to 12 C
4.1.2 Define the Avogadro constant as the number of atoms in 12g of 12 C
4.1.3 Define the mole as the amount of substance containing the number of particles equal to the Avogadro constant

4.2 Molecular and Formula Mass, Molar Mass
4.2.1 Calculate the masses of moles of elements and compounds

- 4.2.2 Calculate the number of moles in a given mass of pure compound
- 4.2.3 Calculate the number of particles (atoms, molecules or formula units) in a given mass of a pure substance

4.3 Types of Solutions
- 4.3.1 Explain the nature of unsaturated, saturated and supersaturated solutions
- 4.3.2 Describe how to distinguish one from the other by addition of a seed crystal
- 4.3.3 State that solutions may be in one of three phases: gases, liquids or solids
- 4.3.4 State that the most common type of solid-in-solid solutions are called alloys, and that the most common type of liquid-in-solid solutions are called amalgams
- 4.3.5 Use the terms miscible and immiscible to describe the solubility of a liquid in a liquid
- 4.3.6 Explain that the miscibility of liquids, and solubility of solids and gases in liquids, depends on intermolecular attractive forces

4.4 Factors Affecting Solubility
- 4.4.1 Define solubility in terms of mass of solute and mass of solvent
- 4.4.2 State that the solubility of solids in liquids depends on the temperature
- 4.4.3 State that the solubility of a gas in a liquid depends on the pressure of the gas above the liquid and temperature
- 4.4.4 Predict whether an ionic compound will dissolve or mix in water using a solubility table

4.5 Concentration of Solutions
- 4.5.1 Concentration
  - 4.5.1.1 Define concentration
  - 4.5.1.2 Define solute, solvent, solution

- 4.5.2 Standard Solutions
  - 4.5.2.1 Define a standard solution as containing a precise mass of solute in a precise volume of solution
  - 4.5.2.2 Calculate molarity, given the mass of solute and volume of solvent
  - 4.5.2.3 Calculate mass, volume or molarity, given any two of these

- 4.5.3 Percent by Mass and PPM
  - 4.5.3.1 Express the concentration of a solution in terms of percent by mass (mass/mass, mass/volume and volume/volume) and parts per million (mass/mass and mass/volume)

- 4.5.4 Ion Concentrations
  - 4.5.4.1 State that ionic substances dissociate into ions when dissolved in water and that some molecular substances ionize when dissolved in water
  - 4.5.4.2 Calculate the concentration of each ion in salt or acid solutions

knowing the concentration of the salt

5.0 Chemical Reactions

5.1 Law of Conservation of Mass

- 5.1.1 State the Law of Conservation of Mass

5.2 Chemical Reactions

- 5.2.1 Writing and Balancing Chemical Equations
  - 5.2.1.1 Define chemical reaction
  - 5.2.1.2 List the changes that indicate a chemical reaction has taken place
  - 5.2.1.3 Define chemical equation
  - 5.2.1.4 Define reactants and products
  - 5.2.1.5 Write chemical equations for simple reactions, given names of reactants and products
  - 5.2.1.6 Use appropriate subscripts to indicate physical states of reactants and products
  - 5.2.1.7 Balance simple reactions, given the name of formulae of the reactants and products

- 5.2.2 Reaction Types
  - 5.2.2.1 Identify a chemical reaction as combination, decomposition, single or double replacement, neutralization, precipitate formation and combustion
  - 5.2.2.2 Write the products for the complete combustion of hydrocarbons
  - 5.2.2.3 Write the product for combination reactions involving:
    - 5.2.2.3.1. two elements
    - 5.2.2.3.2. formation of hydrates
    - 5.2.2.3.3. nonmetal and metal oxide with water

  - 5.2.2.4 Write the products for decomposition reactions of:
    - 5.2.2.4.1. a compound into its elements
    - 5.2.2.4.2. a hydrate
    - 5.2.2.4.3. carbonates

  - 5.2.2.5 Use the Activity Series to determine whether or not single replacement reactions will occur, and write the products of the reaction
  - 5.2.2.6 Write the products for neutralization reactions
  - 5.2.2.7 Predict the formation of a precipitate in a double replacement reaction given solubility rules
  - 5.2.2.8 Predict the products for the addition of acids to carbonates and bicarbonates

- 5.2.3 Net Ionic Reactions
  - 5.2.3.1 Write total ionic equations from molecular equations

5.2.3.2 Identify spectator ions
5.2.3.3 Write net ionic equations from molecular equations

5.2.4 Oxidation and Reduction Reactions
5.2.4.1 Recognize examples of redox reactions that affect materials
5.2.4.2 Define oxidation and reduction in terms of electron transfer
5.2.4.3 Define oxidation number
5.2.4.4 State the range for oxidation numbers is from -4 to +7
5.2.4.5 State that the oxidation number of atoms in a polar covalently bonded molecule depends on the electronegativity values of each atom
5.2.4.6 Deduce the oxidation number of an element in a given compound or ion using the rules for assigning oxidation numbers
5.2.4.7 Define a redox reaction in terms of both change in oxidation number and electron transfer
5.2.4.8 Select examples of redox reactions from a given set of equations
5.2.4.9 Define an oxidizing agent as an element that undergoes reduction
5.2.4.10 Define a reducing agent as an element that undergoes oxidation
5.2.4.11 Identify, in given redox reactions, which reactants are oxidizing agents and which are reducing agents

5.2.5 Stoichiometry
5.2.5.1 Perform mole-mole and mass-mass calculations using a balanced equation
5.2.5.2 Perform calculations involving limiting reagents using a balanced equation
5.2.5.3 Calculate percent yield given experimental yield
5.2.6.4 Explain why experimental yield is different from theoretical yield

6.0 Kinetic Molecular Theory and Gas Laws

6.1 Characteristics and Properties of Gases
6.1.1 List and describe general characteristics of gases
6.1.2 Describe the kinetic theory of gases
6.1.3 Define pressure and list the most common units of pressure
6.1.4 Define Boyle's Law
6.1.5 Use Boyle's Law in calculations
6.1.6 Define and calculate Kelvin temperatures
6.1.7 Define Charles's Law
6.1.8 Use Charles's Law in calculations
6.1.9 Explain the relationship between Boyle's and Charles's Laws
6.1.10 Use the combined gas law in calculations

6.2 Gas Mixtures and Partial Pressure
6.2.1 Explain Dalton's Law of Partial Pressure
6.2.2 Calculate partial pressure of gases

6.2.3 Explain Avogadro's Hypothesis
6.2.4 Define standard temperature and pressure
6.2.5 Define and calculate the molar volume of a gas

6.3 Ideal Gas Law
6.3.1 Define and derive the Ideal Gas Law
6.3.2 Define the universal gas constant
6.3.3 Perform calculations based on gas laws and stoichiometry

EVALUATION:

Quizzes:	40%
Laboratories:	10%
Final Exam:	50%

If a student misses a laboratory session without a *valid documented reason***, a mark of 0 for that lab will be assigned. In order to be eligible to write the final examination (including a supplementary final examination) and pass the course, students must pass (minimum of 50%) the essential laboratory component of the course. A student who misses more than 3 labs without valid documentation will be required to drop the course. Please note that dropping the course without academic prejudice must be done within established College processes and time frames.

** What would be considered as a "valid documented reason" will be at the discretion of the campus administrator in consultation with the faculty responsible for the two Chemistry courses, CH1120 and CH1121.

DATE DEVELOPED:	March 1999	**DATE REVIEWED:**	February 2014
REVISION NUMBER:	2	**DATE REVISED:**	September 2011

Note to instructor: Check PIRS to ensure this outline is the most current version.