THE UNIVERSITY OF NEW SOUTH WALES

SCHOOL OF CHEMISTRY

CHEM1011 Chemistry A
CHEM1031 Higher Chemistry A
and
CHEM1051 Higher Chemistry Medicinal A

INFORMATION FOR STUDENTS

SEMESTER ONE, 2018

Course Coordinators

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SCHOOL OF CHEMISTRY – THE UNIVERSITY OF NEW SOUTH WALES

CHEM1011 –CHEMISTRY A, CHEM1031 – HIGHER CHEMISTRY A, AND CHEM1051 – HIGHER CHEMISTRY MEDICINAL A

INFORMATION FOR STUDENTS

BACKGROUND

Either of the courses CHEM1011 (Chemistry A) or CHEM1031 (Higher Chemistry A) may be taken as the first half of level one chemistry. Chemistry A is suitable for students with limited background in high school chemistry. Higher Chemistry A (available in Semester 1 only) assumes a thorough background in high school chemistry. Either course may be taken as a stand-alone course by students who require only one semester of level one Chemistry. CHEM1051 is available only to students in the Medicinal Chemistry program.

OBJECTIVES OF CHEM1011 AND CHEM1031/CHEM1051

Both Chemistry A and Higher Chemistry A aim to extend your understanding of the basic principles of Chemistry that underlie our understanding of matter by building on your chemical knowledge from High School. After a brief review of the assumed knowledge, these courses cover the structure of atoms and how this relates to the chemical properties of the elements, the nature of chemical bonds, intermolecular forces and the properties of solutions, classes of chemical reactions, and electrochemistry and electrochemical cells. The difference between CHEM1011 and CHEM1031/CHEM1051 is that in CHEM1031/CHEM1051 introductory material will be covered only briefly, allowing more time to for discussion and support of the more complex issues in class time and explore new material in lectures.

REQUIREMENTS

The assumed knowledge for CHEM1011 is one year of senior high school chemistry (*e.g.* Yr 11 in NSW) or equivalent. The School of Chemistry strongly recommends that all students who do not meet this assumed knowledge requirement undertake a bridging course in chemistry or do CHEM1001 – Introductory Chemistry, before entering CHEM1011. Students entering CHEM1031/CHEM1051 should have studied chemistry in Years 11 and 12 (or equivalent) and gained a good grade in their high school chemistry (Band 5 or 6 in HSC, or equivalent).

ASSESSMENT

The formal assessment components of the course and the proportion of your final grade allocated to each is as follows:

Laboratory Assessment	20%	Mid-Semester Test	10%
Online Revision Quizzes	10%	End of Semester Examination	60%

A pass in CHEM1011, CHEM1031 or CHEM1051 requires:

- a total mark of at least 50 for the course,
- a net laboratory attendance of at least 80%,
- all core laboratory skills awarded (*see laboratory manual for more details*), and
- an end of semester examination mark of at least 35.0% (ie: 21.0 out of 60.0).

Supplementary examinations may be offered for academic or other reasons. Students are expected to make themselves available for these should they require them. Averages will ***not*** be given in place of an end of semester or supplementary exam mark. Details of examination time and location can be found on Page 5 of this course pack.

ARRANGEMENT of CLASSES and TIMETABLE

There are 6 hours of class contact per week for the majority of this course. Over the semester there are approximately 36 lectures, 12 tutorial classes and 20 hrs of practical work. The lectures and tutorial classes are organized as follows:

	Lectures (hours/week)	**Tutorial** (hours/week)	**Laboratory** (hours/week)
Week 1	3	–	no lab - do Safety & Ethics online (*see lab manual*)
Weeks 2 – 4	3	1	2
Week 5	2*	1	CHEM1011 = no lab; CHEM1031/51 = 2
Weeks 6	3	1	2
Week 7	3	1	(make–up lab only)
Weeks 8	2*	1	2
Weeks 9 – 12	3	1	2
Week 13	2*	1	(make–up lab only)

*In Week 5, Friday is a Public Holiday. In Week 8, Wednesday is a Public Holiday. As a result, if you have a lecture on either of these days, it will not be running. To compensate, there may be extra lectures held in Week 13. Check your timetable for details.

The times and locations of classes can be found on myUNSW. Students with a timetable clash should see the Chemistry Student Centre (Rooms 104/105 Dalton) to explore alternative arrangements.

CLASS LOCATIONS MAY BE CHANGED IN RESPONSE TO CHANGES IN STUDENT NUMBERS. CHECK YOUR TIMETABLE ON MyUNSW OFTEN NEAR THE START OF SEMESTER.

DIAGNOSTIC TEST

At the start of semester, an online diagnostic test will be available, which will enable you to assess whether your chemistry background is sufficiently strong to allow you to continue in the course you have enrolled in or whether you should transfer to a lower level chemistry course. This test is compulsory, but its result WILL NOT contribute to your assessment. Nevertheless, it is your best interest to do as well as you can so that you can make a realistic decision about which first year chemistry course suits your background.

If you are studying CHEM1011 in Semester 1, and you score a poor mark or a very good mark in the diagnostic test, you could consider changing to a less demanding course (CHEM1001), or the higher level course (CHEM1031). In Semester 2, when CHEM1031 and CHEM1001 are not offered, if you score a poor mark in the diagnostic test, you could consider deferring first year chemistry and taking a chemistry bridging course over summer. If you are considering changing CHEM courses you ***must*** obtain advice from the Chemistry Student Centre before changing your enrolment.

ATTENDANCE

Students are expected to attend all their scheduled classes (lectures, tutorials and laboratories) in all first year chemistry courses. Attendance will be recorded at laboratory and tutorial classes. In particular, a minimum of 80% attendance is required for a satisfactory performance in the laboratory component of the course.

IF YOU MISS MORE THAN ***TWO*** LABORATORY CLASSES YOU WILL AUTOMATICALLY FAIL THE COURSE.

See the Laboratory Manual (in this course pack) for more details on attendance and all other aspects of the laboratory classes.

WORKLOAD

A guide to the workload for this course is that you should spend an hour of independent study for each contact hour. The bulk of the time during semester should be spent preparing tutorial problems and preparing and writing up laboratory reports. Closer to the examination, more time should be spent reviewing lecture material.

STUDENT INFORMATION and COMMUNICATION with STUDENTS

Announcements relevant to the course will be made in lectures. Hence missing lectures, or arriving late, may cause you considerable difficulties. Students should also consult notices posted in the Moodle "Important Announcement' Section. Communication via your student email account (z1234567@unsw.edu.au) will be used for important announcements and late changes to arrangements. It is your responsibility to check your email at least every two days and make sure that your account does not become 'over-quota' (*i.e.* full) or, if you have messages redirected, that the redirection is functioning.

CHECK YOUR UNSW EMAIL AT LEAST EVERY TWO DAYS!

ASSESSMENT COMPONENTS

Laboratory Assessment

This is described in the laboratory manual.

Revision Quizzes

There are online Revision Quizzes most weeks during the semester. The schedule and topics are listed in the schedule inside the cover of the course pack.

- To attempt a quiz, logon to Moodle (https://moodle.telt.unsw.edu.au/) and select the 'Revision Quizzes' section.
- Select the relevant quiz you wish to attempt. In most cases, only one quiz will be available to you at any given time.
- In most cases, quizzes are open for *one* week only. They generally open at 12:01am on Monday morning, and close at 11:59pm on Sunday night. These details are displayed in the quiz itself, before you make an attempt. During this time, you can make up to **three attempts** to answer the quiz questions (using any help you desire, taking as long as you want up until the quiz close time). You will most likely get a different version of the question on each attempt.

- You must get full marks on one of your three possible attempts at each quiz to get 1 added to your total quiz score. If you do not get full marks in any of your attempts, you will be awarded a score of zero for that quiz. Your best 10 quiz marks during the semester will be used to calculate your total quiz score, which will account for 10% of your final overall mark for this course.
- After the period for answering a quiz has closed, you will no longer be able to attempt the quiz. However, you will be given feedback and hints on how to answer the questions in the quiz.
- The Revision Quizzes are primarily designed to be a learning process, not as an assessment process.

Mid Semester Test

A test of 45 minutes' duration will be held near the middle of the semester (see Page 5 for more details). Material to be covered in the mid-semester test will be notified closer to the test time. Information about how mid-semester tests are conducted is provided later in this document. You need to read this information because it explains your responsibilities with regards to these tests.

End Of Semester Examination

The end of semester examination will cover the entire contents of the course. As a guide, the amount of material on a syllabus topic in the exam will be in proportion to the number of lectures for the topic in the syllabus. A number of questions in the final examination will relate directly to the laboratory component of the course. Students who do not perform to a satisfactory level in the end of semester examination may be offered additional assessment (see Page 5 for more details).

The final examinations for CHEM1011 and CHEM1031/CHEM1051 will both be of 2 hours' duration and will contain questions requiring written answers (similar to the tutorial problems) and multiple choice questions.

COURSE COMPONENTS

Lectures

Many announcements relevant to the course will be made in lectures. Furthermore, some lectures may be recorded, but there is no guarantee that the lecture recording technology will capture the class correctly or even at all. Hence missing lectures, or arriving late, may cause you considerable difficulties.

Tutorials

Attendance at tutorial classes is compulsory. The tutorials provide an opportunity for students to consolidate the material of the course. Tutors will explore problems that students are having with the course. The tutorial notes contain a core set of typical problems associated with the course topics. The tutorial problems are those necessary to master the material in the course. To maximise the benefit of tutorials you should attempt all problems in a tutorial set prior to the tutorial. Many questions in the final examination will be very similar in style and content to those in the tutorials.

Laboratory Classes

The laboratory classes provide an opportunity to learn the concepts and practice the calculations presented in lectures. Laboratory classes are also the place to learn practical skills and they are also the place where those skills are assessed and hence they are a compulsory component of all first year chemistry courses.

No students with an unsatisfactory laboratory record (either due to poor laboratory work or to inadequate attendance) will be considered for a pass in the course. You ***must attend*** *at least 80% of the scheduled laboratory classes in the semester. Medical certificates or other documentation do not compensate for absences.*

You must **READ THE INTRODUCTION IN THE LABORATORY MANUAL** to be aware of all the requirements for passing the laboratory component of this course. Here are some of the main points regarding laboratory classes:

- **Safety eyewear must be worn at all times in the laboratory.**
- **A laboratory coat and fully enclosed footwear must be worn in the laboratory. You will not be permitted to work in thongs or open-top shoes or sandals or without a laboratory coat and safety eyewear.**
- The schedule of experiments is inside the front cover of the **course pack**.
- All experiments require pre–lab work to be completed before your lab class.
- You must attend the laboratory class shown on your official timetable.
- You must arrive at the laboratory on time or you will be excluded from the class.
- Repeat students must apply to the First Year Laboratory Coordinator if they want exemption from laboratory classes. Exemption is not automatic and is decided on a case by case basis.

SEE THE LABORATORY MANUAL FOR MORE DETAILS, including what to do if you are unavoidably absent from a lab class, how to submit reports, and the criteria for grading your laboratory work.

NEED HELP?

There are several people who can help you with problems. The appropriate person may differ depending on the problem.

- For problems relating to lectures − see your lecturer.
- For tutorial problems − see your tutor.
- For laboratory problems − see your demonstrator, or alternatively ask your tutor (if time permits) in tutorials.
- For help with answering questions (tutorials problems, lab exercises, Revision Quizzes, etc) − follow the link "*Chemistry Study Help*" in the "Assessment & Course Information" section on Moodle.
- For course related and/or personal difficulties that may be affecting your performance in first year chemistry − see the Chemistry Student Centre, (Room 105, Dalton Building).
- For all other enquires (including special consideration, enrolment issues or Moodle issues) − contact Steve, the Student Support Manager (firstyearchem@unsw.edu.au).

Chemistry Study Help

The School of Chemistry has a study area, located on the Ground floor of the Dalton Building, in the Gibson Computer Lab (G11). You are welcome to use this study area for quiet study any time between 8:30am and 6pm.

For information regarding additional help in your Chemistry studies, log on to Moodle, and follow the link "*Chemistry Study Help*" in the "Assessment & Course Information" section.

RESOURCES

Text Books

- Blackman, A., Bottle, S., Schmid, S., Mocerino, M., and Wille. U, Chemistry, John Wiley and Sons, 3rd Ed 2016.
- Aylward, G.H. and Findlay, T.J.V., SI Chemical Data, (6th ed.), Wiley, 2008 (or later).
- CHEM1011/CHEM1031/CHEM1051 Course Pack, School of Chemistry, UNSW (On sale at the UNSW Bookshop).

MID-SEMESTER EXAMINATIONS and SUPPLEMENTARY EXAMINATIONS CONDUCTED by the CHEMISTRY STUDENT CENTRE

Dates, Time And Location

Mid-Semester Test

The mid-semester test for this course is currently planned for ***Week 6*** during your regular scheduled lecture time on ***Wednesday 11 April***. Please check your timetable to confirm which of these dates and times apply to you. These dates are subject to change based on enrolment numbers and venue availability. Note: you MUST attend your test during the lecture time for which you are officially enrolled. You may not attend a different sitting without prior approval. If you cannot attend this lecture time due to a permitted timetable clash, please contact firstyearchem@unsw.edu.au as soon as possible. ***Permitted clashes do not automatically excuse absences from the test***.

The official details will be posted on Moodle approximately 1–2 weeks before the date of the examination.

- ❐ IT IS ENTIRELY YOUR RESPONSIBILITY TO ASCERTAIN THESE DETAILS.
- ❐ No information regarding time and location will be given over the telephone.

Supplementary Examination (by direct invite only)

The supplementary exam period for this semester is ***Saturday 14 July – Saturday 21 July***. Whilst you may not *plan* on needing this exam, please be aware that you may be required to attend for either special consideration or academic reasons. Therefore, ***all students are expected to make themselves available for this exam period*** should they be required to attend. No alternative dates or times will be guaranteed.

Attending Examinations

You may not be permitted to sit an examination if you do not have your UNSW student I.D. card with you.

- ❒ You should endeavour to be in attendance at the given location well before the "doors open" time.
- ❒ You will be admitted to the room from the time listed as "doors open".
- ❒ You may take in only pens, pencils, **electronic calculator** (listed in the UNSW list of approved calculators) and drawing instruments. No other materials may be taken into the room unless officially notified. If you do bring such material along with you on the day, it must be left outside the examination room.
- ❒ **You should note that neither the Chemistry Student Centre, nor the University, will accept responsibility for any material left outside the examination room. Hence we urge very strongly that you do not bring such items with you.**
  (NB. A pen, a 2B-pencil, and eraser are essential for multi-choice exams.)
- ❒ Once you are inside the room normal University Examination rules will apply.
- ❒ UNSW Examination rules state that you may not wear a watch during your exams. Please leave them in your bag with the alarm off for the duration of the exam. If you do not bring a bag, you may place your watch in a clear resealable plastic bag under your seat before the exam starts.
- ❒ You **may not bring** into the examination room:
  - a mobile phone, music player, mp3 player or iPod, or any other communication device
  - a bag or bags
  - paper, books, etc
  - electronic or discipline-specific dictionaries.

Procedures for Examinations

- ❒ You should take your seat in the room following the instructions given by the supervisor and the invigilators.
- ❒ Immediately you are seated you must place your UNSW student I.D. card, "photo" side up on your desk.
  Your card must remain on your desk throughout the examination, and be visible at all times.
- ❒ When you received your examination paper, you must NOT turn this paper over, nor write anything on it until you are instructed to do so. This includes your name and ID details.
- ❒ When permitted to write, you must enter your ID details, as instructed, on the test paper and on the Generalised Answer Sheet. **Marks will be deducted** if you do not enter both your name and student number as required for authentication and for scanning purposes.
- ❒ Each answer must be recorded on the Generalised Answer Sheet by using pencil to fill in the oval corresponding to your chosen answer. The answers entered on the Generalised Answer Sheet will be used to determine your test mark.
  - **You will not be permitted to leave during this time. If you finish early you must remain seated in your place until the end of the allocated time.**
  - You will not be admitted later than fifteen (15) minutes past the "examination commences" time.
- ❒ If you have any query, etc. raise your hand and keep it raised until one of the invigilators attends to you.
- ❒ A "10 minutes to go" warning will be given.
- ❒ At the end of the allocated time a "pens down - cease writing" command will be given.
  You must cease writing immediately, place your pen on the table, and close your examination paper. Remember that if you do not heed this command, disciplinary action may be taken against you. You must remain seated in your place until instructed to leave by the supervisor.
  NOTE: At this stage, if you have not written your name and student number on your paper, you will NOT be given extra time to do so. It is entirely your own responsibility to make sure your details are on your exam paper BEFORE the "pens down" instruction is given.

Feedback

- ❒ The mark you obtain for the Mid-Semester Test will be returned via Moodle within 10 working days, including the average mark for the test across the entire course.

GENERAL ARRANGEMENTS FOR UNSW EXAMINATIONS

Check the Examination Timetable through the link on the myUNSW website for the location and time of the final examination. Make sure you know where the examination is to be held – it may be in a location you have never attended before (*e.g.* Randwick Racecourse). Allow plenty of time to get to the examination venue on the day.

Medical documentation or other "considerations" for all examinations must be submitted to Student Central (in the Chancellery) within 3 days of the examination. Documentation submitted after the deadline will not be taken into account for the assessment. You should also notify the Chemistry Student Centre in Room 105 Dalton.

Where documentation has been submitted requesting a "consideration" for the end-of-semester examination, because the student's performance may have been affected sufficiently to alter his/her grade in the overall assessment, the student may be required to sit for a supplementary examination consisting of a written paper and (in some rare cases) an oral examination. Oral examinations will mainly be used to resolve cases for students whose results are borderline and whose performance may be disadvantaged in a written paper due to misadventure.

If you have applied for "consideration", you should arrange to make yourself available for possible further assessment. Notification of details of the further assessment will be sent via your student email address (z1234567@unsw.edu.au).

If you have applied for "consideration", and if you were present at the final examination, but fail to be present for the supplementary examination, then your effort at the final examination will be used in determining your overall mark.

Notification of a student’s requirement to attend a supplementary exam may not be possible until after the end of the initial examination period. Students making travel arrangements during the supplementary examination period should take this into account.

Prior to the date of the of the supplementary examination
you must notify the Chemistry Student Centre
whether you will be attending the examination.
This is your responsibility.
The Chemistry Student Centre can be contacted on 9385-4666 during office hours.

ACADEMIC MISCONDUCT

Students and staff are, of course, governed by the normal laws which regulate our everyday lives. But in addition the University has its own code of rules and conduct, and can impose heavy penalties on students who breach them. Penalties range from failure in a subject, loss of privileges, fines, payment of compensation, and suspension, to exclusion from study for a certain period or even permanent expulsion from the University.

It is important to realise, however, that misconduct within the University covers a much wider field than simply behaviour which is offensive or unruly, or which may cause damage to other people or property. Misconduct which may lead to a student being disciplined within the University includes anything regarded as **academic misconduct *according to current academic usage*,** as well as any conduct which impairs the reasonable freedom of other persons to pursue their studies or research or to participate in University life.

It is most important that students realise just how broad the definition of Academic misconduct may be. It certainly covers practices such as cheating or copying or using another person's work. Sometimes, however, practices which may have been acceptable at school are considered to be misconduct according to current Academic usage within a University. For example academic misconduct can occur where you fail to acknowledge adequately the use you have made of ideas or material from other sources (see the UNSW Student Guide for examples).

The following are some of the actions which have resulted in students being found guilty of academic misconduct in recent years:

- impersonation in examinations;
- failing to acknowledge the source of material in an assignment;
- taking of unauthorised materials into an examination;
- submitting work for assessment knowing it to be the work of another person;
- improperly obtaining prior knowledge of an examination paper and using that knowledge in the examination.

Students found guilty of academic misconduct are usually excluded from the University for two years. Because of the circumstances in individual cases, the period of exclusion can range from one semester to permanent exclusion from the University.

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CHEM1011

Lecture content and learning objectives by week

Recommended Texts

- Blackman, A., Bottle, S., Schmid, S., Mocerino, M., and Wille. U, Chemistry, John Wiley and Sons, 3rd Ed 2016.
- Aylward, G.H. and Findlay, T.J.V. SI Chemical Data, (6th ed.) 2008 (or later).
- CHEM1011/CHEM1031/CHEM1051 Course Pack, sold at the UNSW Bookshop.

The topics listed in this syllabus, the exercises in the tutorial sets, and assignments in the laboratory programme define the examinable material. The 'Refs' column lists section numbers (and sometimes page numbers) from the textbook (Blackman, *et al*. 'Chemistry') provide suitable reading material.

Week	Lecture Content CHEM1011	Refs	Learning objectives. (At the end of this section you should be able to…)
1-2	Atoms and their structure. Nomenclature of simple organic and inorganic ions. Stoichiometry and balancing equations Mole Concept. Solution Stoichiometry Empirical vs Molecular formula; Percentage composition by mass Common reagents and their reactions (acids, bases (Bronsted), oxides, carbonates). Gases: properties, diffusion. Ideal gas law; determination of molar mass; density. Gas mixtures: partial pressures, Dalton's law.	1.1 – 1.3, 3.1 – 3.6 11.1 6.1 – 6.6	Name the constituent parts of an atom, together with their relative masses and charges. Calculate numbers of protons, neutrons, electrons in atoms of a particular element. Name simple inorganic compounds and write the formulae for simple compounds from their name. Write and balance simple chemical equations. Calculate molecular weight from chemical formula. Calculate % by mass of each element in a compound and determine empirical formula from % by mass. Calculate concentration of solutions in various units. Calculate yield in a chemical reaction, determine the limiting reagent. Define what is meant by an acid and a base. Know the names and formulae of common acids and bases. Predict the products of reactions of acids and bases with oxides, carbonates etc. Describe the properties which distinguish gases from other states of matter. Calculate properties of gases and gas mixtures using the ideal gas equation.
2-4	Atomic spectroscopy, Hydrogen spectrum, Rydberg equation. Quantum mechanical principles and the quantum mechanical model of the atom. Orbitals and quantum numbers from the Schrödinger equation. Electronic configurations of atoms; Pauli exclusion principle, Hund's rule. Electronic configurations and the periodic table.	4.1 – 4.5 4.6 – 4.9 5.1, 5.2	Calculate wavelength from frequency and *vice versa* for electromagnetic radiation. Use the Rydberg equation to calculate the wavelengths emitted or absorbed by a H atom. Calculate photon energy for e.m. radiation from its frequency. List the allowed values of the quantum numbers for orbitals in hydrogen–like atoms.

Week	Lecture Content CHEM1011	Refs	Learning objectives. (At the end of this section you should be able to…)
	Configurations of ions. Diamagnetic and paramagnetic species. Isoelectronic species. Periodicity of atomic and ionic sizes, ionization energies and electron affinities. Introduction to bonding: electronegativity and bond polarity. Bond types (metallic, ionic, covalent)		Sketch the shapes of *s*, *p*, and *d* orbitals. Write ground–state electron configurations for all main group elements and first–row transition metals and ions of these elements using 'arrows in boxes' and '$1s^1$' notation. Identify isoelectronic species and predict relative sizes of these species. Predict the magnetic properties of isolated atoms and ions. Predict relative sizes, ionization energies, electron affinities and electronegativities of atoms based on position in the periodic table. Relate the difference in electronegativity between two elements to the type of bond formed between them.
4-6	Bonding: Lewis diagrams; electron deficient & expanded valence shell species Lewis Structures Resonance structures; formal charges. Lewis Structures VSEPR theory and molecular shape; polarity of molecules Valence bond theory, hybridized orbitals. Sigma and pi bonds, multiple bonds. Delocalisation.	5.3 – 5.5 5.6	Draw Lewis diagrams for simple molecules. Identify species where the octet rule is violated. Draw multiple Lewis structures for species where the distribution of electrons is ambiguous. Assign formal charges to atoms in a Lewis structure and use these to select the most likely Lewis structure for a species. Use VSEPR theory to predict the shapes of molecules and polyatomic ions. Predict whether molecules are polar or non–polar. Assign hybridized orbitals to the central atoms in molecules and polyatomic ions. Sketch the overlap between orbitals on adjacent atoms which gives rise to bonding. Identify species where extended orbital overlap gives rise to delocalization of electrons.
6-7	Intermolecular forces: dipole–dipole, dispersion forces, hydrogen bonds. Deviations from ideal gas laws; van der Waals equation; *a* and *b* constants. States of matter. Comparison of properties of solids, liquids and gases; Solids - types (atomic, molecular, ionic, metallic, network); crystalline vs amorphous. Liquids - general description; viscosity and surface tension.	6.7, 6.8 7.1 – 7.3 10.1, 10.3 10.5	Identify the types of intermolecular forces present between particular species. Relate the strength of intermolecular forces to properties such as dipole moment and polarizability. Use the van der Waals equation to calculate the pressure of a non–ideal gas. Explain the origin and size of the *a* and *b* constants. Classify solids as ionic, molecular, or metallic based on their properties. Relate viscosity and surface tension to

Week	Lecture Content CHEM1011	Refs	Learning objectives. (At the end of this section you should be able to…)
	Properties of solutions: energy changes on dissolution, solubility, 'like dissolves like', Raoult's law, ideal and non ideal solutions; positive and negative deviations from Raoult's law. Distillation.		intermolecular forces. Understand the process of dissolution at the molecular level. Predict the relative solubility of a substance in a range of solvents.
7	Mid-semester test – 1 lecture slot.		
8-10	Equilibrium: Definition of K; general calculations using K; Le Chatelier's principle. Acid/base (definitions; Lewis, Bronsted), pH, pOH, Kw, pKw; strong acid and base solutions; K_a, K_b, weak acid and base solutions, hydrolysis of salts, buffers.	9.1, 9.2, 9.4, 9.5, 11.2, 11.3	Understand the concept of chemical equilibrium. Apply Le Chatelier's principle to systems at equilibrium. Apply general equilibrium concepts to the specific case of acid-base equilibria. Calculate the pH and pOH of aqueous solutions of strong and weak acids and bases. Recognise which salts will hydrolyse, and calculate the pH of their solutions. Calculate pH of buffer solutions. Calculate the formulation of a buffer solution of specified pH.
10-11	Thermodynamics. System and surroundings. Heat and heat capacity; endothermic and exothermic processes. Internal energy, work and the First law. Enthalpy, q_p, and q_v; internal energy change and enthalpy change. Calorimetry. Standard enthalpies of sublimation, fusion, vaporization. Hess' law, enthalpies of formation, enthalpy changes for chemical reactions, enthalpy of combustion. Enthalpy of atomization, ionization enthalpy; electron affinity; bond energy; lattice enthalpy; Entropy. Entropy changes for phase transitions and chemical reactions. Definition of Gibbs energy, Gibbs energy changes for chemical reactions, relation to equilibrium, temperature dependence of the equilibrium constant K.	8.1-8.8	Perform calorimetric calculations using heat capacities. Perform calculations involving internal energy, heat and work. Be familiar with a range of processes for which standard enthalpies are tabulated. Use Hess' Law in enthalpy change calculations. Calculate standard enthalpy changes using standard enthalpies of formation. Utilize average bond energies in thermochemical calculations. Calculate standard entropy changes using standard entropies. Calculate standard Gibbs energy changes using standard free energies of formation. Calculate standard Gibbs energy changes using standard enthalpies of formation and standard entropies. Use standard Gibbs energy change to predict spontaneity of a reaction at a specified temperature, and to calculate K at a specified temperature.
11-12	Redox reactions and electrochemistry: Definitions (oxidation, reduction, oxidising	12.1-12.8	Identify redox reactions and the species which are oxidized and reduced.

Week	Lecture Content CHEM1011	Refs	Learning objectives. (At the end of this section you should be able to…)
	agent, reducing agent) oxidation numbers, balancing ion-electron half equations. Standard potentials; hydrogen electrode; electrochemical cells; cell diagrams; cell notation; cell potential; Batteries, Nernst equation; cell potentials and equilibrium constants; Electrolysis; corrosion and corrosion prevention. Revision for final examination.		Calculate oxidation numbers for elements in molecules and ions. Balance redox equations using the half–equation method. Know the meanings of the terms: galvanic cell, anode, cathode, salt bridge, emf, standard hydrogen electrode. Represent a galvanic cell using standard notation. Calculate the emf of a cell from standard reduction potentials. Calculate the emf of a cell operating with non–standard concentrations. Relate cell emf to the equilibrium constant. Interpret the chemical processes in corrosion in terms of redox reactions. Use a chemical understanding of corrosion to describe methods for reducing corrosion. Calculate the amount of products formed during electrolysis from the current and duration of electrolysis.

CHEM1031/CHEM1051

Lecture content and learning objectives by week

Recommended Texts

- Blackman, A., Bottle, S., Schmid, S., Mocerino, M., and Wille. U, Chemistry, John Wiley and Sons, 3rd Ed 2016.
- Aylward, G.H. and Findlay, T.J.V. SI Chemical Data, (6th ed.) 2008 (or later).
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The topics listed in this syllabus, the exercises in the tutorial sets, and assignments in the laboratory programme define the examinable material. The 'Refs' column lists section numbers (and sometimes page numbers) from the textbook (Blackman, *et al*. 'Chemistry') provide suitable reading material.

Week	**Lecture Content** **CHEM1031/CHEM1051**	**Refs**	**Learning objectives.** (At the end of this section you should be able to…)
1	<u>Statement of assumed knowledge:</u> It will be assumed that students entering CHEM1031/CHEM1051 will have already mastered their prior chemistry, including: Atoms and their structure. Nomenclature of simple organic and inorganic ions. Stoichiometry and balancing equations. Mole Concept. Solution Stoichiometry Empirical vs Molecular formula; Percentage composition by mass. Common reagents and their reactions (acids, bases (Bronsted), oxides, carbonates). <u>Lectures in Week 1 will mainly be on:</u> Gases: properties, diffusion. Ideal gas law; determination of molar mass; density. Gas mixtures: partial pressures, Dalton's law.	1.1 – 1.3, 3.1 – 3.6 11.1 6.1 – 6.6	<u>From assumed knowledge you should be able to:</u> Name the constituent parts of an atom, together with their relative masses and charges. Calculate numbers of protons, neutrons, electrons in atoms of a particular element. Name simple inorganic compounds and write the formulae for simple compounds from their name. Write and balance simple chemical equations. Calculate molecular weight from chemical formula. Calculate % by mass of each element in a compound and determine empirical formula from % by mass. Calculate concentration of solutions in various units. Calculate yield in a chemical reaction, determine the limiting reagent. Define what is meant by an acid and a base. Know the names and formulae of common acids and bases. Predict the products of reactions of acids and bases with oxides, carbonates etc. <u>At the end of Week 1 you should also be able to:</u> Describe the properties which distinguish gases from other states of matter. Calculate rates of effusion and diffusion of gases. Calculate properties of gases and gas mixtures using the ideal gas equation.
2-3	Atomic spectroscopy, Hydrogen spectrum, Rydberg equation. Quantum mechanical principles and the quantum mechanical model of the atom. Orbitals and quantum numbers from the Schrödinger equation.	4.1 – 4.5 4.6 – 4.9 5.1, 5.2	Calculate wavelength from frequency and *vice versa* for electromagnetic radiation. Use the Rydberg equation to calculate the wavelengths emitted or absorbed by a H atom. Calculate photon energy for e.m. radiation from its frequency.

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	Electronic configurations of atoms; Pauli exclusion principle, Hund's rule. Electronic configurations and the periodic table. Configurations of ions. Diamagnetic and paramagnetic species. Isoelectronic species. Periodicity of atomic and ionic sizes, ionization energies and electron affinities. Introduction to bonding: electronegativity and bond polarity. Bond types (metallic, ionic, covalent)		List the allowed values of the quantum numbers for orbitals in hydrogen–like atoms. Sketch the shapes of *s*, *p*, and *d* orbitals. Write ground–state electron configurations for all main group elements and first–row transition metals and ions of these elements using 'arrows in boxes' and '$1s^1$' notation. Identify isoelectronic species and predict relative sizes of these species. Predict the magnetic properties of isolated atoms and ions. Predict relative sizes, ionization energies, electron affinities and electronegativities of atoms based on position in the periodic table. Relate the difference in electronegativity between two elements to the type of bond formed between them.
4-6	Bonding: Lewis diagrams; electron deficient & expanded valence shell species Lewis Structures Resonance structures; formal charges. Lewis Structures VSEPR theory and molecular shape; polarity of molecules Valence bond theory, hybridized orbitals. Sigma and pi bonds, multiple bonds. Delocalisation. Molecular orbital (MO) theory, bonding and antibonding MOs, stability of simple diatomic species, bond order. Allocating electrons in MO energy level diagrams. Second-row element diatomic species, effect of *s* and *p* orbital mixing on energies of MOs, and electron filling patterns.	5.3 – 5.5 5.6 5.7	Draw Lewis diagrams for simple molecules. Identify species where the octet rule is violated. Draw multiple Lewis structures for species where the distribution of electrons is ambiguous. Assign formal charges to atoms in a Lewis structure and use these to select the most likely Lewis structure for a species. Use VSEPR theory to predict the shapes of molecules and polyatomic ions. Predict whether molecules are polar or non–polar. Assign hybridized orbitals to the central atoms in molecules and polyatomic ions. Sketch the overlap between orbitals on adjacent atoms which gives rise to bonding. Identify species where extended orbital overlap gives rise to delocalization of electrons. Identify the modes of atomic orbital interactions that give rise to bonding and antibonding MOs. Draw MO energy level diagrams, add symbols for electrons, and calculate bond order.
6-7	Intermolecular forces: dipole–dipole, dispersion forces, hydrogen bonds. Deviations from ideal gas laws; van der Waals equation; *a* and *b* constants.	6.7, 6.8	Identify the types of intermolecular forces present between particular species. Relate the strength of intermolecular forces to properties such as dipole moment and polarizability. Use the van der Waals equation to calculate the

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	States of matter. Comparison of properties of solids, liquids and gases; Solids - types (atomic, molecular, ionic, metallic, network); crystalline vs amorphous. Liquids - general description; viscosity and surface tension. Properties of solutions: energy changes on dissolution, solubility, 'like dissolves like', Raoult's law, ideal and non ideal solutions; positive and negative deviations from Raoult's law. Distillation. Surfactants, molecular membranes, micelles.	7.1 – 7.3 10.1, 10.3 10.5	pressure of a non–ideal gas. Explain the origin and size of the *a* and *b* constants. Classify solids as ionic, molecular, or metallic based on their properties. Relate viscosity and surface tension to intermolecular forces. Understand the process of dissolution at the molecular level. Predict the relative solubility of a substance in a range of solvents. Apply Raoult's law to calculate vapour pressure of mixtures of volatile solvents; calculate the composition of vapour of such mixtures. Relate surfactant structure to roles in micelle and membrane stability.
8-10	Equilibrium: Definition of K; general calculations using K; Le Chatelier's principle. Acid/base (definitions; Lewis, Bronsted), pH, pOH, Kw, pKw; strong acid and base solutions; K_a, K_b, weak acid and base solutions, hydrolysis of salts, buffers.	9.1, 9.2, 9.4, 9.5, 11.2, 11.3	Understand the concept of chemical equilibrium. Apply Le Chatelier's principle to systems at equilibrium. Apply general equilibrium concepts to the specific case of acid-base equilibria. Calculate the pH and pOH of aqueous solutions of strong and weak acids and bases. Recognise which salts will hydrolyse, and calculate the pH of their solutions. Calculate pH of buffer solutions. Calculate the formulation of a buffer solution of specified pH.
10-11	Thermodynamics. System and surroundings. Heat and heat capacity; endothermic and exothermic processes. Internal energy, work and the First law. Enthalpy, q_p, and q_v; internal energy change and enthalpy change. Calorimetry. Standard enthalpies of sublimation, fusion, vaporization. Hess' law, enthalpies of formation, enthalpy changes for chemical reactions, enthalpy of combustion. Enthalpy of atomization, ionization enthalpy; electron affinity; bond energy; lattice enthalpy; Born-Haber cycle. Entropy. Entropy changes for phase transitions and chemical reactions. Definition of Gibbs energy, Gibbs energy changes for chemical reactions, relation to	8.1-8.8	Perform calorimetric calculations using heat capacities. Perform calculations involving internal energy, heat and work. Be familiar with a range of processes for which standard enthalpies are tabulated. Use Hess' Law in enthalpy change calculations. Calculate standard enthalpy changes using standard enthalpies of formation. Utilize average bond energies in thermochemical calculations. Calculate standard entropy changes using standard entropies. Calculate standard Gibbs energy changes using

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	equilibrium, temperature dependence of the equilibrium constant *K*. Thermodynamics and biology, ATP/ADP.		standard free energies of formation. Calculate standard Gibbs energy changes using standard enthalpies of formation and standard entropies. Use standard Gibbs energy change to predict spontaneity of a reaction at a specified temperature, and to calculate *K* at a specified temperature.
11-12	Redox reactions and electrochemistry: Definitions (oxidation, reduction, oxidising agent, reducing agent) oxidation numbers, balancing ion-electron half equations. Standard potentials; hydrogen electrode; electrochemical cells; cell diagrams; cell notation; cell potential; Batteries, Nernst equation; cell potentials and equilibrium constants; Electrolysis; corrosion and corrosion prevention. Important applications of electrochemistry: pH electrodes, electron transfer chain in respiration, photosynthesis. Revision for final examination.	12.1-12.8	Identify redox reactions and the species which are oxidized and reduced. Calculate oxidation numbers for elements in molecules and ions. Balance redox equations using the half–equation method. Know the meanings of the terms: galvanic cell, anode, cathode, salt bridge, emf, standard hydrogen electrode. Represent a galvanic cell using standard notation. Calculate the emf of a cell from standard reduction potentials. Calculate the emf of a cell operating with non–standard concentrations. Relate cell emf to the equilibrium constant. Interpret the chemical processes in corrosion in terms of redox reactions. Use a chemical understanding of corrosion to describe methods for reducing corrosion. Calculate the amount of products formed during electrolysis from the current and duration of electrolysis.