

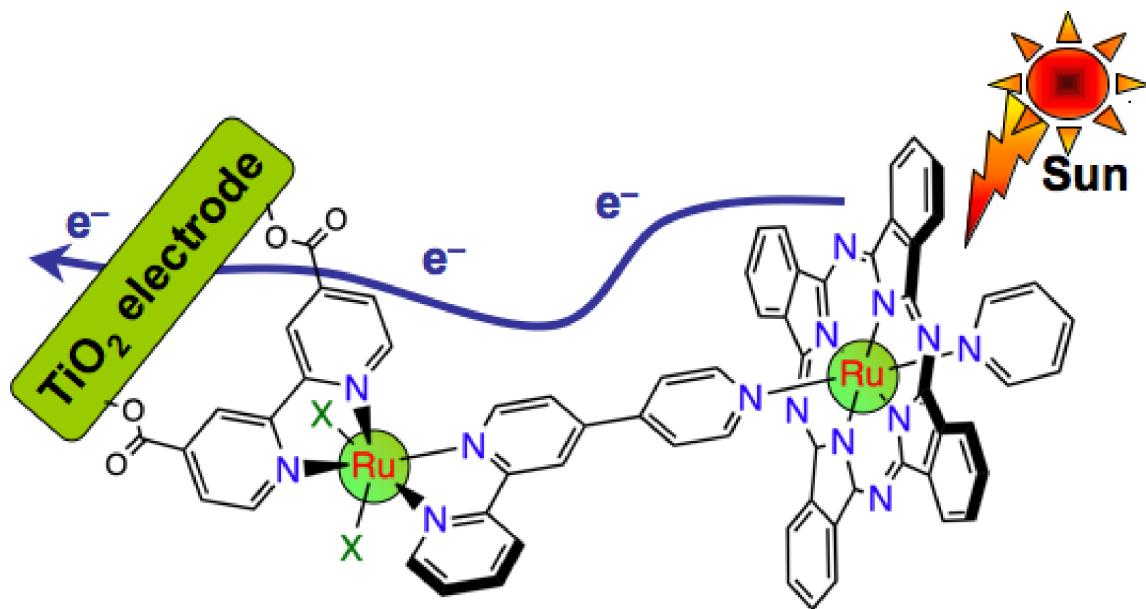
THE UNIVERSITY OF NEW SOUTH WALES

SCHOOL OF CHEMISTRY

Chemistry A: Atoms, Molecules, and Energy

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

LABORATORY MANUAL



Complete the details below at your first laboratory class.

Family name: Other names:

Day/time of laboratory class:

Laboratory (room 133 or 165): Workspace:

Demonstrator:

(You must know your demonstrator's name in order to submit reports for marking.)

Recent Winners of the Nobel Prize in Chemistry

2017	Joachim Frank, Richard Henderson, Jacques Dubochet	Development of cryo-electron microscopy for the high-resolution structure determination of biomolecules in solution.
2016	Jean-Pierre Sauvage, J. Fraser Stoddart, Ben Feringa	Design and synthesis of molecular machines.
2015	Tomas Lindahl, Paul Modrich and Aziz Sancar	Mechanistic studies of DNA repair.
2014	Eric Betzig, Stefan W. Hell and William E. Moerner	The development of super-resolved fluorescence microscopy.
2013	Martin Karplus, Michael Levitt, and Arieh Warshel	The development of multiscale models for complex chemical systems.
2012	Robert J. Lefkowitz and Brian K. Kobilka	Studies of G-protein-coupled receptors.
2011	Dan Shechtman	The discovery of quasicrystals.
2010	Richard F. Heck, Ei-ichi Negishi, and Akira Suzuki	Developing new, more efficient ways of linking carbon atoms together using metal-based catalysts.
2009	Venkatraman Ramakrishnan, Thomas A. Steitz, Ada E. Yonath	Studies on the structure and function of the ribosome (the producer of proteins in cells using DNA as a template).
2008	Osamu Shimomura, Martin Chalfie, Roger Y. Tsien	Discovery and development of the green fluorescent protein, GFP.
2007	Gerhard Ertl	Studies of chemical processes on solid surfaces.
2006	Roger D. Kornberg	Studies of the molecular basis of eukaryotic transcription.
2005	Yves Chauvin, Robert H. Grubbs, Richard R. Schrock	Development of the metathesis method in organic synthesis.
2004	Aaron Ciechanover, Avram Hershko, Irwin Rose	Discovery of ubiquitin-mediated protein degradation.
2003	Peter Agre, Roderick MacKinnon	Discoveries concerning channels in cell membranes.
2002	John B. Fenn, Koichi Tanaka, Kurt Wüthrich	Development of methods for identification and structure analyses of biological macromolecules; (Fenn and Tanaka) development of soft desorption ionisation methods for mass spectrometric analyses of biological macromolecules; (Wüthrich) development of nuclear magnetic resonance spectroscopy for determining the three-dimensional structure of biological macromolecules in solution.
2001	William S. Knowles, Ryoji Noyori, K. Barry Sharpless	(Knowles and Noyori) work on chirally catalysed hydrogenation reactions; (Sharpless) work on chirally catalysed oxidation reactions.
2000	Alan Heeger, Alan G. MacDiarmid, Hideki Shirakawa	Discovery and development of conductive polymers.
1999	Ahmed Zewail	Studies of the transition states of chemical reactions using femtosecond spectroscopy.
1998	Walter Kohn, John Pople	(Kohn) development of the density-functional theory; (Pople) development of computational methods in quantum chemistry.
1997	Paul D. Boyer, John E. Walker, Jens C. Skou	(Boyer and Walker) elucidation of the enzymatic mechanism underlying the synthesis of adenosine triphosphate (ATP); (Skou) discovery of an ion-transporting enzyme, Na^+, K^+ -ATPase.
1996	Robert F. Curl Jr., Sir Harold Kroto, Richard E. Smalley	Discovery of fullerenes.
1995	Paul J. Crutzen, Mario J. Molina, F. Sherwood Rowland	Work in atmospheric chemistry, particularly concerning the formation and decomposition of ozone.

Chemistry is a broad science and it finds application in many fields. Some winners of the Nobel Prize in Chemistry are not 'chemists' in the traditional sense, but their knowledge of chemistry enabled them to make the discoveries which earned them a Nobel Prize.

CHEM1011/CHEM1031/CHEM1051

LABORATORY MANUAL



You must attend 80% of laboratory classes and acquire all core skills in order to pass the laboratory component. See the 'Assessment' section in the 'Information for Students' earlier in the course pack, and the 'Assessment of Laboratory Work' section of the Laboratory Manual.



READ THE SCHEDULE INSIDE THE FRONT COVER OF THE COURSE PACK to see which experiment is to be performed each week. The list below may not be in the order that you will do the experiments.

Title	Page
The Laboratory Course	4
Assessment of Laboratory Work	5
Submission of Laboratory Reports.....	6
Absences from Laboratory Classes	7
Ethics and Plagiarism	8
Safety and Good Practice in the Laboratory	9

EXPERIMENTS

Safety and Ethics in the Chemical Laboratory	11 (complete before 1 st lab)
Finding Your Way around the Lab	15
Preparation of a Standard Solution.....	21
Volumetric Analysis – Standardisation of a Base	25
Volumetric Analysis – Back Titration	33 SUBMIT REPORT
Corrosion [CHEM1031 only].....	41
Intermolecular Forces.....	49 SUBMIT REPORT
Chemical Equilibrium	59 SUBMIT REPORT
Acid–Base Titration	71
Buffers	79
Thermochemistry	87 SUBMIT REPORT
Galvanic Cells	101 SUBMIT REPORT

Techniques

Weighing by difference on an analytical balance	22
Use of a volumetric flask.....	23
Use of a pipette.....	27
Use of a burette.....	28
Using a Digital Multimeter	43
Using a pH meter.....	73
Drawing scientific graphs	73

The First Year Laboratory coordinator would like to acknowledge the valuable suggestions made by many demonstrators, including Dr Joe Brophy, Diana Gershwin, Conrad Gillard, Dr Rima Raffoul-Khoury, and Tom Yu.

THE LABORATORY COURSE

Welcome to the laboratory course for Chemistry A: Atoms, Molecules and Energy (CHEM1011/CHEM1031/CHEM1051). We hope that you enjoy your time in the chemistry laboratory and learn skills that you can use in your future career, both at university and beyond. Please read the following material carefully before your first lab class, it contains detailed information on the expectations, assessments, and outcomes of the laboratory course.

FAILURE TO READ THE FOLLOWING MATERIAL WILL NOT BE ACCEPTED AS A REASON FOR NOT COMPLETING LABORATORY ASSESSMENT TASKS.

Before your first laboratory class

You **MUST** complete the 'Safety and Ethics in the Laboratory' exercise (later in this manual) before your first laboratory class and you must also purchase your safety glasses and lab coat. **Safety eyewear, enclosed footwear, and laboratory coats are compulsory.** Safety eyewear meeting Standards Australia guidelines can be purchased from the Optometry Clinic located in the Rupert Myers Building between 10–12 and 2–4 each weekday. **Regular spectacles CANNOT be used as safety eyewear and regular safety glasses cannot be worn over spectacles.** If you wear spectacles you must wear safety eyewear in the laboratory, either in the form of overglasses, safety goggles, or prescription safety glasses. Disposable paper laboratory coats are not permitted.

Before each laboratory class

You must complete the required pre-lab work and watch any associated skills videos for that experiment. The pre-lab work usually consists of gathering and writing up the safety information for chemicals that you will be using in the laboratory from a web page (see link in Moodle under the 'Laboratories/Practicals' heading), and may also contain questions for you to answer or calculations for you to do, to prepare you for the next experiment.

You must bring the completed pre-lab work along with your laboratory manual (or at least the section of the laboratory manual for that week's experiment) to each laboratory class.

YOU WILL NOT BE ALLOWED TO START AN EXPERIMENT IF YOU HAVE NOT COMPLETED THE PRE-LAB WORK.

You **MUST** bring to each laboratory class the appropriate safety equipment (safety eyewear, lab coat), and be appropriately dressed. Your footwear must cover the entire foot; 'ballet flats', thongs or other open footwear is not permitted. You should attempt to wear non-synthetic fibres in the laboratory, and ideally your clothes should cover as much of you as possible.

YOU WILL NOT BE PERMITTED TO WORK IN THE LABORATORY WITHOUT PROPERLY FITTING SAFETY EYEWEAR, A LABORATORY COAT, AND ENCLOSED FOOTWEAR.

Please be punctual – laboratory classes are only 2 hours long. Your demonstrator or the laboratory supervisor will often make announcements at the start of the class, if you are late you will miss these and you will then be disadvantaged. Your attendance will be recorded; a roll will be taken 10 minutes into and 10 minutes from the end of the laboratory class, and you must be present at both roll calls to qualify for 'satisfactory attendance' in the laboratory class.

IF YOU ARRIVE MORE THAN 20 MINUTES LATE TO A LABORATORY CLASS YOU WILL BE REFUSED ENTRY AND MARKED ABSENT.

In the laboratory

You will be **assigned a laboratory workspace** in your first laboratory class; you must **always work at that workspace**.

Your top priority is to **always work safely**. Keep in mind the hazards of the substances and equipment you will be using – you will have learned about these in your preparation for the lab class. If you have any doubt about a safety issue ask your demonstrator, the lab supervisor or the lab staff. Keep your lab coat and safety eyewear on at all times in the lab.

You should always **record results and observations in a professional manner**. This means recording all numerical observations to the correct precision (determined by the equipment) i.e., with the appropriate number of significant figures. All results must be recorded in indelible pen; correction fluid and pencil are not permitted. If you make a mistake, simply draw a single line through the error before writing the correct value or observation. When leaving your workspace to record a result (e.g., at a balance) you must take your results sheet with you and record the result onto it. Trying to remember the results or writing on your hand is unprofessional, and if you do this you will be penalised.

No later than ten minutes before the end of the class (the time is shown on monitors in the lab) you must cease experimental work to allow sufficient time for you to clean your workspace and equipment so that it is in a fit state for the next student. This will be checked by your demonstrator and failure to do this will result in loss of marks. You may also use the time to complete your laboratory write-up.

Change of laboratory class

Staff will not change laboratory classes for students after the deadline for students to change their timetable in myUNSW, unless the student has documentation from the Student Equity and Disabilities Unit requiring them to take laboratory classes at a particular time, or has an otherwise intractable timetable clash. Note in particular that staff will not change laboratory class times for students in response to requests relating to external commitments, especially employment.

Lab exemption for repeating students

Exemption from laboratory classes for repeating students MAY be granted, but students MUST apply via email to the First Year Laboratory coordinator for exemption before the end of week 4 of semester. To be eligible for exemption from laboratory classes students must have acquired all core skills and satisfied attendance requirements in the previous 12 months. Students who have not done lab classes in the previous 12 months or have been previously exempt from lab classes will not be granted exemption. Lab exemption will not be granted if the course has changed substantially since the last time a student attempted it.

ASSESSMENT OF LABORATORY WORK

Throughout the laboratory classes you will be assessed by your demonstrator on your competency in certain skills (see table below). This assessment will be done both in real-time in the laboratory (for manual lab skills) and retrospectively based on your written reports. The table below lists all of the skills assessed in the laboratory work for this course.

Core (required) skills	Non-core (graded) skills
Analytical glassware	Applying chemical principles
Chemical glassware	Chemical calculations
Ethics induction	Chemical equations
Fume cupboard	Describing chemical changes
Heating materials	Experimental accuracy
Safety induction	Mastery (Applying chemical principles)
Scientific graphing	Mastery (Experimental accuracy)
Titration	Professionalism
Weighing	Recording observations
	Safety awareness
	Time management

All of the assessable skills for each experiment are listed in the lab manual, including a checklist of *criteria* which tell you what you have to do to be awarded each skill – see the 'Feedback' panel in the results or report section of each experiment. Criteria which apply throughout an experiment or report (e.g. 'correct states of matter' on chemical equations) must be **completely** met (i.e. not even one error) for you to be awarded that criterion in that experiment. Make sure you look at the Feedback panel as part of your pre-lab preparation

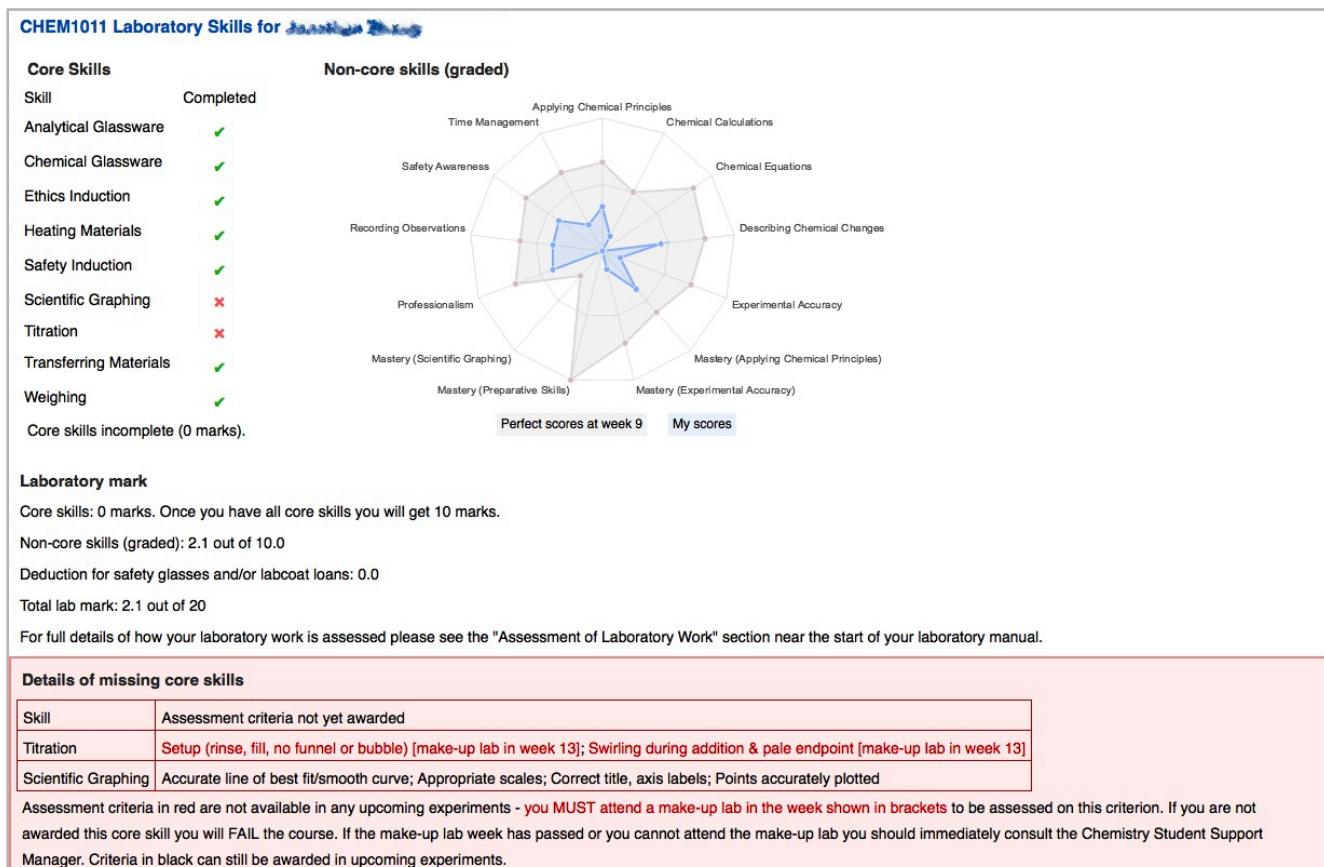
By the end of the semester you **must** be awarded **every** core skill. Once you obtain every core skill you will automatically gain half of the marks allocated to lab work in the overall course assessment and thus pass the lab component of your course. **However, if you have not been awarded every core skill you will fail the laboratory component of the course (and thus the course itself).** There will be multiple opportunities for you to gain each core skill allowing you to gain competency in the core skills at your own pace. Make-up lab classes, although primarily for students who have been absent, provide a last chance for you to acquire core skills that you could not demonstrate in the regular lab classes.

The non-core (graded) skills allow you to gain marks in addition to those awarded once you acquire all the core skills. Each time you completely meet the criteria for a non-core skill you will be awarded marks for that skill. If you demonstrate competency in all the non-core skills in all of the experiments that they are listed in you would gain the other half of the

laboratory marks (in addition to the half for core skills). 'Mastery' skills are awarded for exceptional performance, e.g., having a titre within 1% of the accepted value, and carry higher marks than other non-core skills.

Look at the skills list in the lab manual for the experiment you will be doing so you know what skills you can gain in each lab class. Some of the skills have an associated video (link on the pre-lab page) that shows a first year student demonstrating aspects of the skill. Please **watch** these videos before each laboratory class as part of your preparation for the laboratory class. Your demonstrator will also demonstrate skills in the lab class, but if you watch the video before the class you will progress more quickly through the experiment and likely gain more marks in the non-core (graded) skills.

A link on the pre-lab web page takes you to a report of your progress in acquiring skills, similar to that shown below. You will be able to see which core skills you have obtained (and are yet to obtain), and the marks you have gained for non-core skills, compared to the maximum a student could have obtained at the current week of the semester, displayed on a 'radar' (or 'web' or 'spider') chart. Radar charts are used in many industries to monitor the professional skills of their staff, and identify weaker areas for development. Please allow a week from each lab class for your demonstrator to update your online skills record.



Your end of semester laboratory mark will be made up of a core skills component (half of the maximum laboratory mark) and a non-core (graded) skills component (up to the remaining half of the maximum laboratory mark) minus any deductions for having to borrow a labcoat or safety eyewear.

SUBMISSION OF LABORATORY REPORTS

To provide you with feedback on the quality of your calculations, equations etc. your demonstrator will mark some of your laboratory reports out of the lab class. These reports are indicated in the schedule and table of contents with 'SUBMIT REPORT' listed next to the experiment. For all other experiments your demonstrator will check your report in the lab as you complete the experiment or will ask you to bring the completed report to the following lab class where they will check that you have completed the report, and assess the quality of the relevant non-core (graded) skills. Many reports can be completed in the laboratory class and it is in your best interests to do so.

Reports submitted for marking must be physically submitted to your demonstrator's mailbox during business hours. Emailed photographs of reports are not acceptable. You must submit the actual report from the laboratory class showing the results you recorded in the laboratory. A photocopy or rewritten copy of your report is not acceptable. All results must be

recorded in **indelible PEN (blue or black)** directly into the manual. **Pencil (or erasable pen) is only permitted for graphical work.** White-out or liquid paper is not permitted.

Reports that are required to be submitted must be submitted according to the following schedule:

Monday lab classes	submit report by 5.00 pm Thursday of the same week.
Tuesday lab classes	submit report by 5.00 pm Friday of the same week.
Wednesday lab classes	submit report by 10.00 am Monday of the next week in semester.
Thursday lab classes	submit report by 5.00 pm Monday of the next week in semester.
Friday lab classes	submit report by 5.00 pm Tuesday of the next week in semester.

If the due date/time for a report falls on a public holiday, the deadline moves to the next working day. Laboratory reports must be submitted into your demonstrator's mailbox located **outside room 131 in the Chemical Sciences Building**. This is adjacent to the lift foyer near the first year chemistry labs. School of Chemistry buildings are not open to undergraduate students on weekends, public holidays, or outside of business hours.

If you are ill on the day that a report is due, please email firstyearchem@unsw.edu.au within 24 hours of the report deadline with a copy of your medical certificate. We will advise you of a new deadline, but in most cases this will be 5pm on the next working day following the expiry of your medical certificate.

Reports submitted after the deadline will be corrected to provide feedback, but will **not** earn any non-core skills for the report.

ABSENCES FROM LABORATORY CLASSES & MISSED CORE SKILLS

If you miss a laboratory class for **any reason** you will be permitted to do a make-up lab. Students who fail to be awarded core skills should also attend a make up lab to get a last chance to be awarded their missing core skills. Make-up labs are available **only in the weeks listed as 'make-up lab' in the schedule inside the front cover of the Course Pack**. You may not do a make-up lab in any other week. A make-up lab compensates for an absence from any one lab class, including those where submitted reports are required. Presentation of a medical certificate does **not** cancel an absence. If you are absent for an experiment you must do the experiment at the next make-up lab, not any subsequent make-up lab. For example if you were absent in week 3 and there are make-up labs in weeks 7 and 13, you must attend the week 7 make-up lab. The week 13 make-up lab only allows students to repeat experiments scheduled for weeks after the first make-up lab.

1. Register online by following the link near the top of the web page where you get your pre-lab safety information **before 5pm on the Thursday before the make-up lab week**. Several make-up times will be scheduled in the make-up lab week for each course, however not all regular lab times will have a matching make-up lab. Choose a make-up lab which fits your timetable. Make-up labs which do not have sufficient students may be cancelled, and in this case the Chemistry Student Centre will contact affected students by email to their official UNSW email address. Please **check your UNSW email daily**.
2. Do the pre-lab exercise for the experiment that you missed or revise the pre-lab for an experiment where you failed in previous attempts to gain core skills.
3. Attend the lab at the scheduled time with safety glasses, lab coat, enclosed footwear, and the completed safety pre-lab work. Show your student card to the demonstrator to confirm your identity.
4. The demonstrator supervising the make-up lab will allocate you to a workspace and supply an orange-coloured form.
5. Fill in only your part of the orange form. The demonstrator will fill in their part, and give you the bottom tear-off slip.
6. Carry out the experiment, recording results and doing calculations as you would for a regular lab class. The make-up lab demonstrator will assess your work to the same standards as your regular demonstrator.
7. Staple the orange slip to the report/results sheet for the lab. This serves as your record that you completed a make-up lab. If the experiment you have repeated is a 'submit report' experiment you must submit the report, with the orange slip stapled to the front, to your demonstrator's mailbox by the deadline listed above in the submission of reports section (based on your make-up lab day, not your regular lab day).

NOTE: you can attend only one make-up lab class in each make-up lab week, and you can do **only one experiment in a make-up lab class**. You can only attend a make-up lab if you have been absent or have not been awarded core skills. You are NOT permitted to do a make-up lab to have a second chance at acquiring non-core (graded) skills.

ETHICS AND PLAGIARISM

It is unethical to present the work of others as your own. The School of Chemistry takes breaches of ethical behaviour very seriously. The School Policy on Ethical Conduct is available on your course website. You MUST read this policy. Most experiments will be performed individually but some experiments, or parts of experiments, will be carried out in collaboration with other students, and these are clearly labelled as such.

Full UNSW Student Conduct policy: <<https://student.unsw.edu.au/conduct>>

UNSW Learning Centre guidelines on writing scientific reports: <<https://student.unsw.edu.au/writing>>

For avoiding plagiarism in reports, consult: <<https://student.unsw.edu.au/plagiarism>> (URLs verified 21 January, 2018)

Individual report / assignment

- The submission is all your own work, except where clearly acknowledged
- No sections are copied from another student (including consulting a photographed copy of another student's report)
- The report must be an original, not a photocopy or rewritten copy of any other report
- All material directly quoted from another source (textbook, lab manual, web or other source) is in quotation marks or otherwise distinguished from your own writing AND is fully referenced.
- All material, information or ideas summarized from other sources are fully referenced.
- You have not given your report to another student to enable them to copy from it.
- All observations are honestly presented and not invented or altered to fit a preconceived outcome
- You have included a signed coversheet with your assignment

Group report / assignment – additional requirements

- All contributors to the group work are acknowledged
- Any additional guidelines specific to the particular group work are followed.

SAFETY & GOOD PRACTICE IN THE LABORATORY

The rules listed below apply all to UNSW Chemistry laboratories, including the laboratory in which you will be working.
Read this section carefully!

General Rules

1. Laboratory work may be carried out **only** during allocated class times.
2. Experimental work is only to be performed on the exercise designated for that particular class. **Unauthorised experiments are prohibited.**

Safety

3. **Safety spectacles (Australian Standard) must be worn at all times in the laboratory.** If you wear prescription spectacles you must wear over-glasses covering your spectacles while you are in the laboratory or obtain prescription safety glasses. Overglasses and safety goggles purchased from the School of Optometry are professionally fitted.
4. **Sensible clothing** must be worn in the laboratory, including a **laboratory coat**. Disposable paper laboratory coats are not permitted – they present a fire hazard and do not provide any protection from corrosive substances.
5. **Fully enclosed shoes must be worn in the laboratory.** Students will **not** be permitted to work in thongs or open sandals or shoes which do not cover the upper surface of the feet. High heeled shoes and 'ballet flats' are not permitted. Bare feet are prohibited throughout the Chemistry Building because of the danger of cuts from broken glass and from spilled chemicals.

THE LAWS REGARDING SAFETY CANNOT BE SUSPENDED JUST BECAUSE YOU ARRIVE AT THE LABORATORY WITHOUT SAFETY GLASSES, ENCLOSED FOOTWARE, OR YOUR LABORATORY COAT.

IF YOU ARRIVE AT THE LABORATORY WITHOUT ONE OR MORE OF THESE ITEMS YOU WILL BE TOLD TO LEAVE THE LABORATORY AND YOU WILL BE MARKED ABSENT FOR THAT CLASS.

The Students of Chemistry Society (SOCS) has lockers on the ground floor of the Dalton Building which you can hire for a semester. If you keep your safety glasses and lab coat in a locker you will not risk being excluded from a lab class because you left them at home. See the staff in the School of Chemistry Administration area to get in contact with the SOCS committee or look for advertisements on the noticeboards near the start of semester.

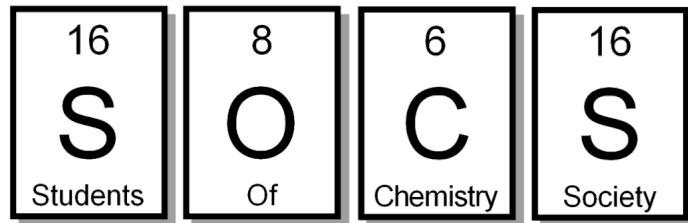
6. **Fume cupboards** must be used where reactions produce toxic or unpleasant gases.
7. Eating and drinking are not permitted in the laboratory. Wash your hands thoroughly after each lab class.
8. Read and be aware of the contents of the notes on safety in the laboratory (green pages included with the lab. manual).

Good Practice

9. Take care not to contaminate reagents. In particular:
 - (a) **DO NOT** mix or interchange stoppers and lids of bottles. Always replace the lid or stopper on a bottle after you use it.
 - (b) **DO NOT** insert squash pipettes into reagent bottles. Pour a small amount of the reagent into a suitable vessel and use the squash pipette on this sample.
 - (c) **Never** return any chemical or solution to the reagent bottle. Discard unused portions.

Housekeeping

10. Place *waste material* in the appropriate containers – solids in the solid waste box, broken glass in the glass container, organic liquids in the drums in the fume cupboard, aqueous solutions (other than those containing heavy metal ions) down the drain, flushed with plenty of water. Aqueous solutions of heavy metal ions must not be poured down the drain, but must be poured into the designated container.
11. Please be *economical* with reagents and materials, including distilled water. Only take sufficient amounts of materials for the experiment.
12. Please clean all glassware with detergent provided on the side bench and rinse all glassware thoroughly with water before returning it to the locker or to the service room.



If you study chemistry, SOCS is the society for you! SOCS caters for all students, from first year undergraduate to final year PhD. All students studying chemistry are automatically members. We run a number of social events through the year, as well as helping with school events.

President Aaron Kennedy

Treasurer Karin Schaffarczyk McHale

Secretary Jonathon Ryan

Activities Coordinator Tim Elton

Merchandising Officer Nicole Richardson

ARC Delegate Stephen Bortolussi

Undergraduate Representatives:

Johanna Wordsworth (Nano)
Matthew Taylor (Higher Chem)

Social Events – Semester 1

Trivia Night (April 10th, Week 6)

Think you're hot stuff down at your local? Come and test your trivia knowledge, at a fun night of questions you'll never need to know the answers to again!

International Food Night (Week 11)

Come and experience the diverse culinary skills of the UNSW chemistry department. We are looking for food from all corners of the globe, so if you are interested in contributing please let us know closer to the date.

Chemball (June 1st, Week 13)

The biggest event on the chemistry social calendar, get up in your finest and come let your hair down! Also an opportunity to mingle with the staff getting a little bit silly.

All dates are to be confirmed, so keep an eye out for more details and more events during the semester.

Lockers

SOCS hires out lockers to students, providing them with a place to stash their junk (*i.e.* textbooks, lab coat, safety glasses). Only \$45/year, or \$30/semester! If you have any questions or would like more information, please email chemlockersunsw@gmail.com. Get in quick as they sell out fast!

Contacts

To stay up to date with SOCS events, or to contact the SOCS exec:

FB: UNSW- Students of Chemistry Society

Email: unsw.socs@gmail.com

Web: <http://www.chemistry.unsw.edu.au/current-students/undergraduate/socs>

SAFETY AND ETHICS IN THE CHEMICAL LABORATORY

Your name: _____ Lab day/time: _____ Workspace: _____



This exercise MUST be completed before your first laboratory class. Failure to do so will result in exclusion from the laboratory until you complete this exercise.

Safety

Aim

To gain an understanding of safe practices in a chemical laboratory.

Method

The following questions should be answered as you view the 'Safety Tutorial' on the web. Click the 'Link to Safety and Pre-lab Information' under the 'Laboratories/Practicals' heading in Moodle to get to the 'First Year Chemistry Safety and Pre-lab Information' web page and then click the 'Safety Tutorial' link near the top of that page.

Answer the questions below as you read through the material on the web pages. WRITE YOUR ANSWERS IN PEN. Pencil is not accepted in any laboratory report, except for plotting data points on graphs. Please click the 'Save and Exit' button at the bottom of the web page when you have finished.

1. Look at the photograph of a poor laboratory workspace. Identify as many features as possible which represent poor or dangerous laboratory technique. (You do not need to fill in ALL the spaces below.)

1	
2	
3	
4	
5	
6	
7	
8	
9	
10	

2. State three ways in which you could be exposed to harmful substances in a chemical laboratory and indicate for each, one way in which the risk can be reduced.

Type of Exposure	Risk Reduced By ...

3. What should you check on a fume cupboard before you use it?

4. After you have used a fume cupboard what should you do with the sash?

5. The WorkSafe Australia technical report "List of Designated Hazardous Substances" contains a list of substances which have been classified as hazardous, together with the risk and safety phrases which should appear on the label of the chemical. Write down the appropriate risk phrases and safety phrases which correspond to the codes shown below. **Click on the words 'Risk' and 'Safety' on the ADG label to display the phrase codes.**

Phrase Code	Associated Phrase
Risk Code R36/37/38	
Safety Code S24	

6. It is not only in the chemical laboratory that dangerous substances will be encountered. Many are found at home. Complete the following table by giving an example of a substance which may be found in the home which belongs to each of the ADG classes given below.

ADG Class	Example
3	
5	

7. Describe the dangerous goods (ADG) diamond symbols for flammable and toxic substances, including the label colour, text, and graphics symbols present.

Type of Compound	Description of Dangerous Goods Diamond
Flammable liquid	
Toxic	

8. What are the two GHS signal words used to describe hazardous substances?

9. Describe the GHS pictograms for corrosive substances and for substances which are toxic in natural waterways, including graphics symbols present.

Type of Compound	Description of Dangerous Goods Diamond
Corrosive	
Toxic in natural waterways	

10. What is the text for each of the following GHS hazard and precautionary statements?

Hazard or Precautionary Phrase Number	Hazard or Precautionary Phrase
H226	
H300	
P262	
P273	

11. What must you do before you use any laboratory glassware?

12. What must you do before you use a Bunsen burner?

13. Where do you dispose of used gloves?

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Core Skills (these skills – only chance to be assessed)

Fume cupboard: completed induction questions

Safety induction completed safety questions

Comments: _____

Demonstrator signature and date: _____

Ethics

Answer the questions below after reading the School of Chemistry anti-plagiarism document, available through Moodle. You should also read the information about academic misconduct in the Information for Students section earlier in this Course Pack.

1. List the four means of committing misconduct in academic works as indicated in the School of Chemistry document.
2. In the School of Chemistry Policy document several scenarios are given. Scenario 2 outlines potential problems associated with group work. Outline acceptable steps towards preparing an individual report from a communal set of results.
3. A checklist is provided in the Policy notes for preparing a report. List the four items on the checklist that you consider the most important. If you consider more than four to be equally important you can list them all.
4. If found guilty of Misconduct by the University, list actions that the University may take against you.

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)**Core Skills (these skills – only chance to be assessed)**

Ethics induction: completed ethics questions

Comments: _____

Demonstrator signature and date: _____

FINDING YOUR WAY AROUND THE LAB

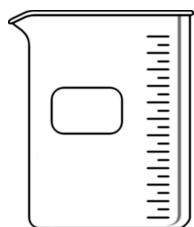
Introduction

The first hour of your first laboratory timeslot in Chemistry is designed to familiarise you with the lab. During this time you will:

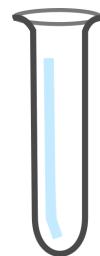
1. Complete your laboratory registration by filling in your lab card.
2. Complete a table showing the names of common laboratory glassware and other equipment.
3. Complete a map of safety and other items in the laboratory.
4. Complete a table showing the locations of common chemicals and other items you will need in later exercises.

Laboratory Glassware and Equipment

Under each piece of glassware or equipment write its name. If you do not know any of the names ask the staff.



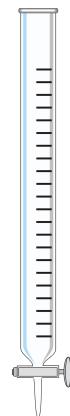
Name:



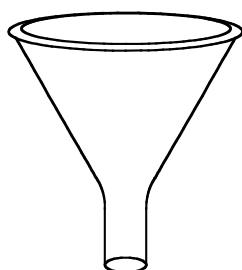
Name:



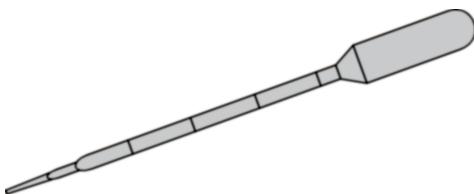
Name:



Name:



Name:

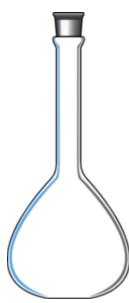


Name:

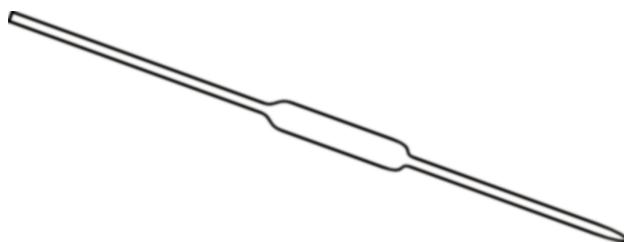


Name (current):

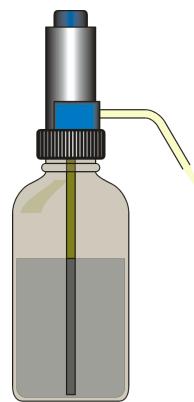
(older name):



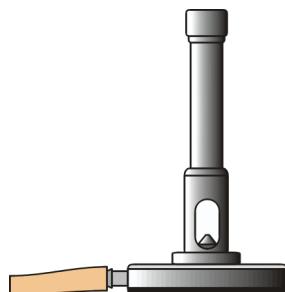
Name:



Name:



Name:



Name:



Name:



Name:

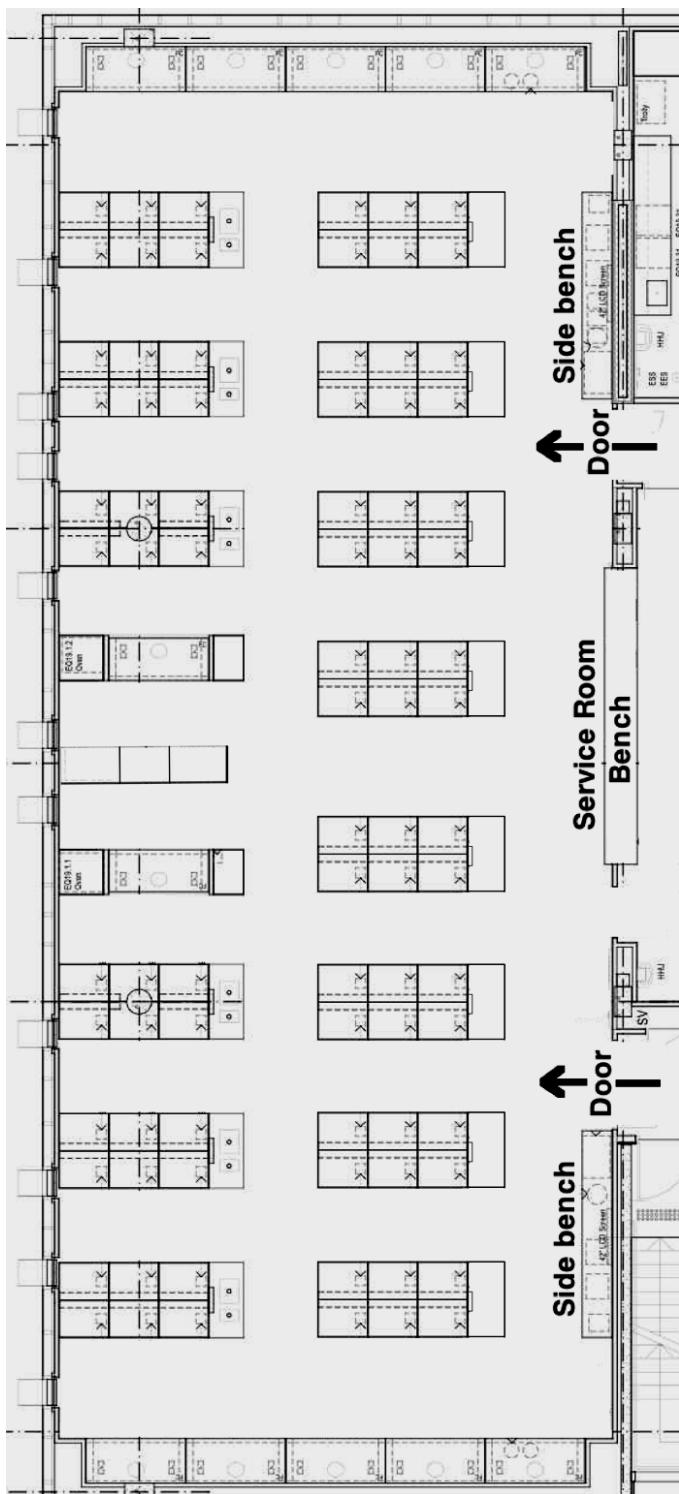


Name:

Safety Items in the Laboratory

Label the map only for the laboratory you are working in, EITHER room 133 OR room 165.

Map for room 133

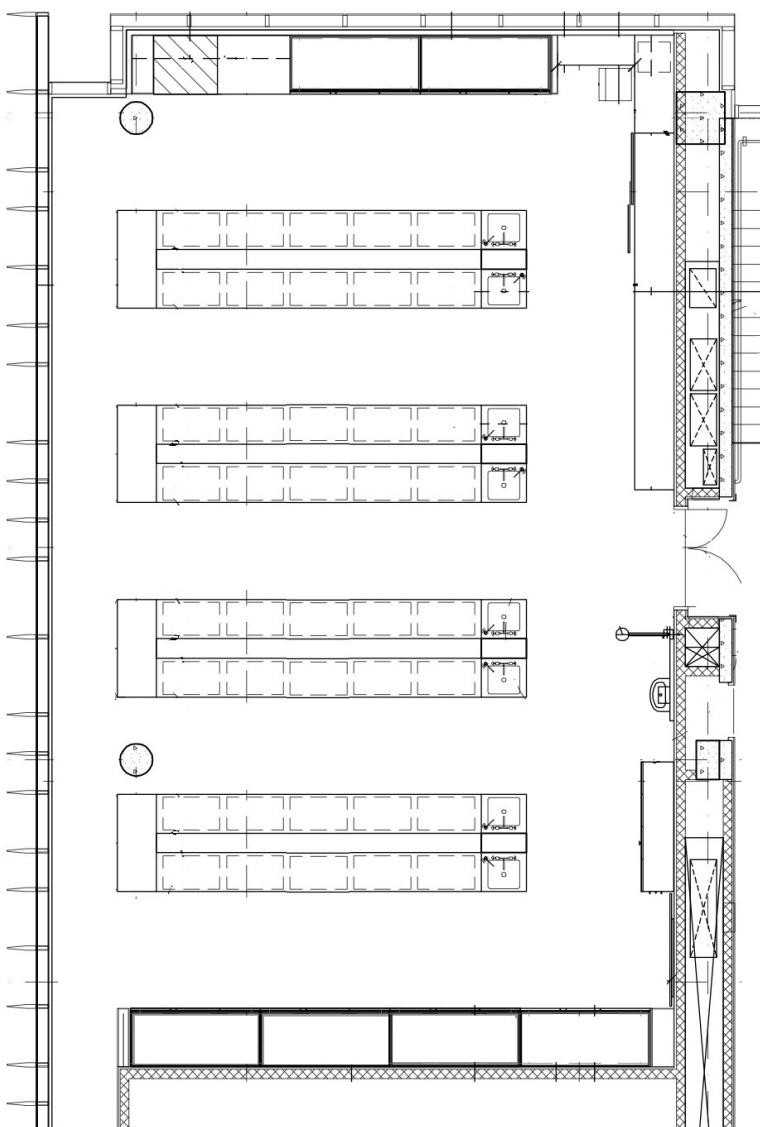


On the map of the laboratory mark the location of each of the following items in the first year laboratory:

- (a) safety showers
- (b) eye washes
- (c) fire extinguishers
- (d) fire blankets
- (e) first-aid kits
- (f) fume cupboards
- (g) disposal containers for organic liquids
- (h) disposal containers for broken glass
- (i) disposal container for used gloves
- (j) disposal containers for other solids
- (k) disposal container for squash pipettes

Also, mark the position ('me') of your workspace on the map.

Map for room 165



On the map of the laboratory mark the location of each of the following items in the first year laboratory:

- (a) safety showers
- (b) eye washes
- (c) fire extinguishers
- (d) fire blankets
- (e) first-aid kits
- (f) fume cupboards
- (g) disposal containers for organic liquids
- (h) disposal containers for broken glass
- (i) disposal container for used gloves
- (j) disposal containers for other solids
- (k) disposal container for squash pipettes

Also, mark the position ('me') of your workspace on the map.

Finding Chemicals, Glassware and Equipment

Next to each chemical or item of equipment write down where you would find it in the laboratory. Possible answers include: my equipment locker, bench-top chemical racks, side benches, bench-ends, issue room, fume cupboard, under fume cupboard, first aid locker. Look in these places to see what they contain.

Item	Location in the laboratory
100 or 150 mL beaker	
test tubes	
test tube rack	
top-loading balance	
analytical balance	
burette	
burette clamp	
2 M ammonia solution	
Bunsen burner	
organic waste container	
steam bath (for heating)	

1. Write the name of the most appropriate piece (or pieces) of equipment or glassware to perform the following tasks.

- Quickly measuring 55 mL of water to within 1 mL.
- Delivering 10.0 mL of a solution into a beaker
- Weighing 12.0 ± 0.2 g of sodium chloride
- Weighing about 0.2 g of sodium chloride to ± 0.001 g
- Temporarily holding waste solutions for later disposal in a sink

2. What do you do with a plastic squash pipette once you have used it?

3. What colour are these taps on the lab benches?

gas vacuum re-circulated water (NOT a general-purpose water tap)

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

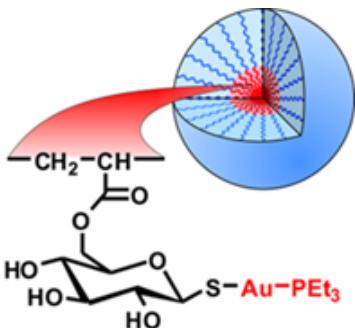
Core Skills (these skills – only chance to be assessed)

Chemical glassware: identifying chemistry glassware

Safety induction: located laboratory safety equipment

Demonstrator signature and date: _____

Cool Research in the School of Chemistry



Prof. Martina Stenzel is studying plastic nanoparticles to hunt out cancer

The delivery of drugs can be improved by packaging the drug into nanoparticles. Nanoparticles for drug delivery typically have sizes below 100 nm and can be prepared from various materials including polymers. In our group we synthesize various polymers to create core-shell nanoparticles – the core holds the drug, often an anti-cancer drug, while the shell makes the particle soluble and determines how it interacts with cells. We use a range of materials starting from synthetic polymers, which we combine with nature's building blocks such as polysaccharides, sugars, peptides, and proteins to create hybrid particles. These particles are then loaded with various anti-cancer drugs. Our projects range from the delivery of metal-based drugs such as cisplatin to peptide drugs or DNA.

In our group we work on many aspects of this technique starting from organic synthesis to polymer nanoparticle preparation to testing these particles on cancer cell lines. We also meet with clinicians to discuss their drug delivery problems.

For more information about this research see:

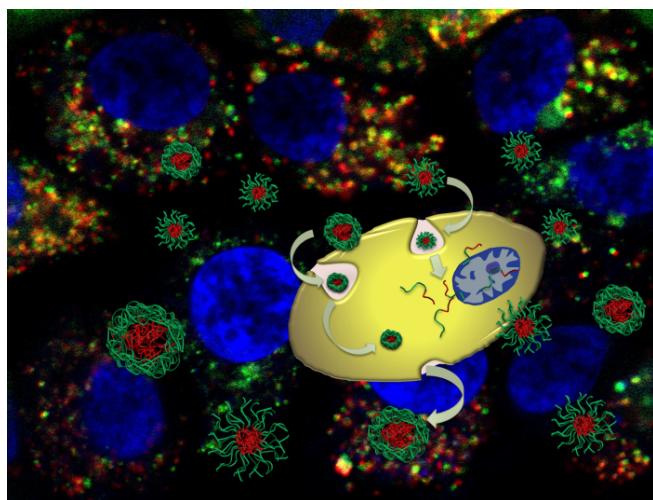
<http://www.chemistry.unsw.edu.au/research/research-groups/stenzel-group>

or contact: Prof Martina Stenzel

9385 4656

Dalton building, room 226

M.Stenzel@unsw.edu.au



PREPARATION OF A STANDARD SOLUTION

Aim

In this experiment you will learn how to accurately weigh out a compound, and then make a solution of known concentration (which will be checked by your demonstrator).

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

Standard solutions are solutions where the concentration of the solute is accurately known. One measure of concentration is molarity, defined as the number of moles of a compound per litre of solution. To make up a standard solution of known molarity a pure compound (the solute) is accurately weighed out before being dissolved in sufficient solvent (usually water) to produce an accurately known volume of solution.

The compound used in this experiment is a copper EDTA complex ($[\text{Cu}(\text{O}_2\text{CCH}_2)_2\text{NCH}_2\text{CH}_2\text{N}(\text{CH}_2\text{CO}_2\text{Na})_2] \cdot 2\text{H}_2\text{O}$) where EDTA is an acronym for ethylenediaminetetraacetic acid. To make up a solution to an accurately known volume, a **volumetric flask** is used. When carefully filled to the mark on the neck of the flask, the volumetric flask used in this experiment will contain 100.0 mL.

There are several ways of accurately assessing the concentration of a solution. In this experiment your demonstrator will determine the concentration of your standard solution using a spectrophotometer (a common chemical technique encountered in later chemistry courses), and determine how accurate your preparation of the standard solution has been.



Preparing a standard solution is all about accuracy. You must measure and record the precise mass of solute and use the volumetric flask correctly (not filling above the mark). Careless technique resulting in inaccurate measurements will produce a solution whose calculated concentration does not equal its actual concentration and you will have to repeat the experiment to be accredited for the skills being assessed.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before coming to the laboratory you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below. The GHS 'signal' words are either 'warning', 'danger', or blank (for substances no significant hazards).

The web page where you get your safety information has a link to VIDEOS about techniques you will use in this experiment. **You should watch these videos before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
ethylenediaminetetraacetic acid copper(II) disodium salt		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Pre-lab work (compulsory, must be completed to be allowed to start the experiment)

What advantage does weighing by difference have? (answer is in the pre-lab video)

Calculate the concentration of a solution made by dissolving 0.600 g of the ethylenediaminetetraacetic acid copper(II) disodium salt (molar mass = 397.74 g mol⁻¹) in water to give a final volume of 100.0 mL.

$$\text{Amount (mol)} = \text{mass} / \text{molar mass}$$

and

$$\text{Molarity} = \text{amount (mol)} / \text{volume (litres)}$$

$$\text{Concentration / mol L}^{-1} =$$

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

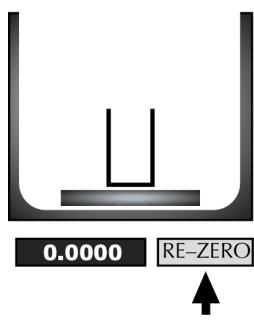
Demonstrator's signature and date

Techniques

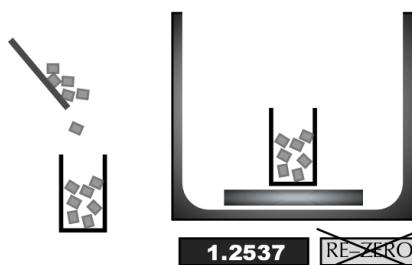
Read the instructions below and watch the videos online (links on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these techniques.

Technique – Weighing by difference on an Analytical Balance

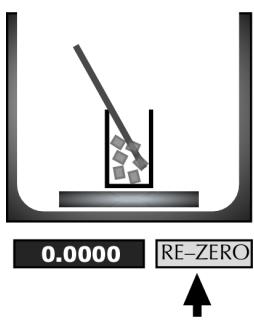
There are many models of balance in the first year laboratory. All balances have a 'tare' button which causes the balance to show zero mass, regardless of what is on it. However the tare button may be labelled differently on different balances. The current balances in the lab have their tare button labelled as 'RE-ZERO', 'O/T', or 'T'.



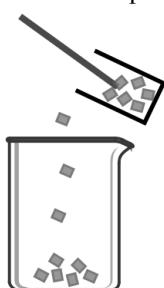
- 1a. If there is a sample tube containing sufficient solid available you can skip to step 2, otherwise place an empty sample tube on balance and re-zero the balance.



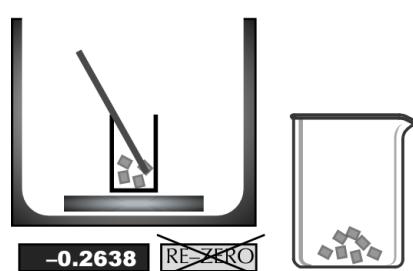
- 1b. Take the sample tube off the balance and add solid to the tube until you have more than the quantity you need to weigh. Check the mass of solid by placing the tube on the balance.



2. Put the spatula into the sample tube (which is on the balance pan). Re-zero the balance.



3. Remove the sample tube from the balance and transfer solid into the vessel where you will be dissolving it. Make sure that any solid removed from the sample tube goes into the vessel.



4. Replace the sample tube and spatula on the balance pan. The balance will read a negative amount. This indicates the mass removed from the sample tube and thus added to the vessel. Repeat 3 and 4 until you have the desired mass in the vessel and record the mass.

Technique – Use of a Volumetric Flask

A volumetric flask is used to prepare an accurately known volume of a solution. The volumetric flasks used in this course are of 250.0 ± 0.3 mL and 100.0 ± 0.15 mL capacity.

Before using a volumetric flask, rinse it with the solvent it is to contain (most often water). A volumetric flask does not need to be dry before use – it simply needs to be clean.

When preparing a solution of a solid in a volumetric flask, dissolve all of the solid in water in a beaker first (possibly with warming on a steam bath if needed, then cooling) and then, without loss, transfer the solution into the volumetric flask through a filter funnel. Rinse the beaker and funnel thoroughly with a stream of water from a wash bottle to ensure no loss of material.

Add water (from the tap, or a beaker) to the volumetric flask to within about 1 cm of the graduation mark on the flask. Using a squash pipette, add more water slowly to the flask until the bottom of the meniscus is in line with the graduation mark.

Stopper the flask and mix the contents by inverting and swirling several times. Hold the stopper in place when swirling to prevent it coming out or solution leaking around the stopper.



Materials Needed

Ethylenediaminetetraacetic acid -copper(II)	100.0 mL volumetric flask	analytical balance
disodium salt	plastic funnel	plastic cuvette
100 mL conical flask	spatula	spectrophotometers

Method

Prepare the ethylenediaminetetraacetic acid copper(II) disodium salt standard solution

Note: In steps 1 and 2 you will use the technique of *weighing by difference* to measure the exact mass of the sample of the ethylenediaminetetraacetic acid copper(II) disodium salt. This technique is described in the 'Techniques' section above and is available to watch as a chemistry skills video online (link available in Moodle).

1. Fill one third of a sample tube with the ethylenediaminetetraacetic acid copper(II) disodium salt. Place a plastic spatula in the sample tube with the ethylenediaminetetraacetic acid copper(II) disodium salt and place on an **analytical balance**. Tare the balance so that it reads zero.

Warning: After carrying out step 1, **do not place the spatula on the bench**, as this may result in a loss of sample, and hence an incorrect determination of the mass of the sample.

2. Using the spatula, transfer without loss (*i.e.* don't spill any), some of the solid from the sample tube into a 100 mL conical flask. Replace the spatula and reweigh the sample tube and spatula. Repeat until you have **about 0.6 g** in your conical flask. Record on the results page the exact mass of the ethylenediaminetetraacetic acid copper(II) disodium salt weighed. The balance reading will be negative since the tube contains less solid for the second reading than the first.
3. Add about 50 mL of water to the conical flask to dissolve the ethylenediaminetetraacetic acid copper(II) disodium salt. Carefully swirl the mixture to aid dissolution.

Warning: Make sure you use the correct technique for using a volumetric flask in the steps below.

4. Using a plastic funnel, transfer the solution into a clean **100 mL** volumetric flask being careful not to lose any. Rinse the conical flask and funnel with a jet of water from a wash bottle several times to ensure **all** the ethylenediaminetetraacetic acid copper(II) disodium salt is transferred to the volumetric flask.
5. Add water carefully from a beaker until the water level is 1 to 2 cm below the graduation mark on the neck of the flask.
6. **Using a squash pipette**, add water drop by drop until the bottom of the meniscus is in line with the graduation mark on the flask (**warning:** if the bottom of the meniscus goes over the graduation mark, you have to start again). Stopper the flask and **mix thoroughly** (this is important!). Hold the stopper in place, turn the flask upside down and swirl.
7. Calculate the concentration of your ethylenediaminetetraacetic acid copper(II) disodium salt solution, based on the actual mass weighed out.

Preparing a sample of the standard solution for analysis

8. Rinse out a **cuvette** using some of the standard solution that you have just made (a new clean squash pipette should be used for the transfer). Repeat, rinsing out the cuvette a second time. Next fill the cuvette to $\frac{3}{4}$ full with the standard solution and take this to your demonstrator.
9. The demonstrator will accurately determine the concentration of the solution and accredit you with the relevant skills.
10. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.

RESULTS

Your name: _____ Lab day/time: _____ Workspace: _____

Preparation of standard ethylenediaminetetraacetic acid copper(II) disodium salt solution

Mass of ethylenediaminetetraacetic acid copper(II) disodium salt (± 0.002 g.)	
--	--

Molar mass of ethylenediaminetetraacetic acid copper(II) disodium salt = 397.74 g mol⁻¹.

Show your calculations here (with appropriate number of significant digits and units in final result)

$$\text{Molarity} = \frac{\text{amount (mol) of solute}}{\text{volume of solution (litres)}}$$

Concentration of ethylenediaminetetraacetic acid copper(II) disodium salt =

Demonstrator's use only	Absorbance: _____	Signature _____
-------------------------	-------------------	-----------------

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Core Skills

Analytical glassware: volumetric flask - correct use

Weighing: analytical balance - weighing by difference

Non-core Skills

Chemical calculations: correct arithmetic correct method correct sig. fig. on results correct units

Experimental accuracy: absorbance within 10% of expected value

Mastery (Experimental accuracy): absorbance within 1% of expected value

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Comments: _____

Demonstrator signature and date: _____

VOLUMETRIC ANALYSIS – STANDARDISATION OF A BASE

Aim

In this experiment you will accurately determine the concentration of a solution of sodium hydroxide by titration with a primary standard. This experiment illustrates the basic techniques and principles associated with volumetric analysis.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

The aim of volumetric analysis is to determine the concentration of one solution by the addition of a stoichiometric amount of a second solution from a burette. If the volume of each of the two solutions used is known accurately, and the concentration of one is also known accurately, the concentration of the other solution can be determined.

The first step in volumetric analysis is to prepare a solution of known concentration. This can be done by accurately weighing an amount of a pure substance and dissolving it in a solvent (often water) to produce a solution of accurately known volume and hence accurately known concentration. You have made a standard solution in a previous experiment. In this experiment the standardised solution will be provided for you, this will allow you to focus on practising your pipette and titration techniques.

To perform a titration, an accurately known volume of one of the solutions (either the standard solution or the solution of unknown concentration) is transferred into an Erlenmeyer (conical) flask using a **pipette**. The second solution is dispensed from a **burette**, which can accurately deliver volumes of the **titrant**, usually with a limit of reading of 0.05 mL.

In order to determine when stoichiometric amounts of the reactants have been mixed (that is, when the **equivalence point** has been reached) an indicator is often employed. In acid/base titrations, the indicator is a dye which changes colour over a specific pH range. At the **equivalence point** of an acid–base titration, a sudden change in acidity of the solution occurs and this is indicated by a colour change of the indicator (the **endpoint** is when the colour change is seen).

A knowledge of the two volumes used and the concentration of one of the solutions allows the concentration of the other solution to be calculated.

The **standardised solution** used in this experiment is a solution of HCl. As mentioned this will be provided to you in this experiment – make sure that you record the concentration in the results sheet as it will be needed for the calculations in the write up.

25.00 mL of the HCl solution is pipetted into a conical flask and titrated with the sodium hydroxide. The amount (mol) of HCl used is known accurately, and since the HCl and sodium hydroxide react in a 1:1 ratio, the amount (mol) of sodium hydroxide consumed is also accurately known. From the average titre, the concentration of the sodium hydroxide can be calculated.

There are two ways of assessing the success of any quantitative chemical analysis. The *precision* of your analysis refers to how close together repeated measurements are. If the method has been carried out very precisely then repeating the analysis will give results very close to the previous results. The *accuracy* of the analysis refers to how close the result you obtain is to the true result. You should be aware that certain factors may affect the accuracy of a titration without degrading its precision. For example, if you consistently overshoot the endpoint (add more titrant than is required to reach the endpoint) then your results may be very precise, but they will certainly not be accurate.



Quantitative chemical analysis must be performed precisely and accurately to be useful. In all titration experiments you will be graded according to the precision and accuracy of your results. Inaccurate and imprecise results will result in a low mark, or you may be asked to repeat a titration.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below. The GHS 'signal' words are either 'warning', 'danger', or blank (for substances no significant hazards).

The web page where you get your safety information has a link to VIDEOS about several of the techniques you will use in this experiment. **You should watch these videos before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
0.1 M HCl		
0.1 M NaOH		
Phenolphthalein		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Pre-lab calculation (compulsory, must be completed to be allowed to start the experiment)

A 25.0 mL sample of a NaOH solution was titrated with 0.125 mol L⁻¹ HCl. The equivalence point was at 24.40 mL of HCl. Calculate the concentration of a NaOH solution.

$$\text{Molarity} = \text{amount (mol)} / \text{volume (litres)}$$

$$\text{Concentration / mol L}^{-1} =$$

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Techniques

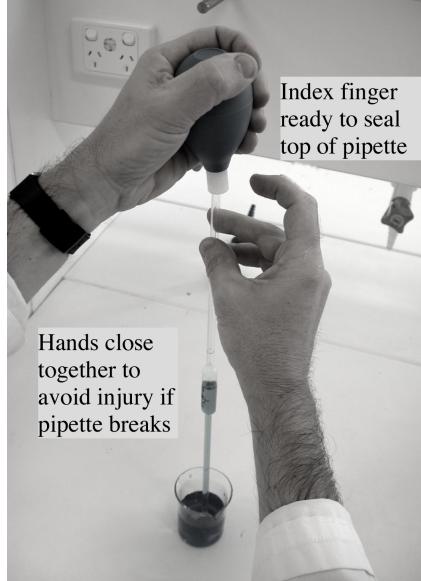
Read the instructions below and watch the videos online (links on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these techniques.

Technique – Use of a Pipette

A **volumetric or transfer pipette** is used in quantitative analysis to transfer an exact single volume of a liquid to a receiving vessel. The glass and plastic types commonly used in the first year laboratories have a single graduation mark at 25.00 ± 0.04 mL above a bulb that stores the bulk of the liquid.

A rubber bulb must always be used to draw liquid into the pipette. **Never use your mouth!** The first time you draw a liquid into a pipette you must use that liquid to rinse the pipette. Only then can you use the pipette to dispense the liquid.

1. Squeeze the bulb to expel air and hold the bulb compressed.
2. Hold the pipette close to the top, above the graduation mark. Gently insert the pipette into the filler. Do not insert it more than 3 or 4 mm into the filler. Do not force the pipette into the filler! **The seal is provided by the pressure you apply when holding the filler and the pipette together.**
3. With the pipette vertical, dip the tip into the liquid to be pipetted and draw up 3 or 4 mL of liquid into the pipette to be used to rinse the pipette. Release your squeeze on the bulb while keeping the seal by applying pressure as you hold the filler and the pipette together. You are not filling the pipette, just take up a little liquid to rinse it.
4. Remove the pipette filler and quickly seal the top of the pipette with your **index finger**. Rotate the pipette until it is horizontal and then rotate it about its length to wet and rinse the interior of the pipette. Discard the rinse liquid.
5. Repeat the rinsing procedure from steps 2 – 4 one more time using your solution.
6. Repeat the filling process from step 3 above, but this time draw the solution well above the graduation mark, but **do not get liquid into the pipette filler**.
7. Wipe carefully down the stem and around the pipette tip with a tissue to remove **excess** liquid.
8. Remove the filler from the pipette and quickly seal the top of the pipette with your index finger. Let the liquid run out until it reaches the mark.
9. Position the pipette into the receiving vessel. Make sure that the tip of the pipette touches the side of the container. Raise your finger from the top of the pipette, until liquid stops flowing. Ensure that the tip of the pipette is always touching the side of the vessel. When drained, remove the pipette from the vessel **but do not blow out the remaining drops.**



Technique – Use of a Burette

A burette is used to titrate (add) a solution of one reactant (the titrant) contained in a burette to a measured volume of a solution of another reactant in a conical flask.

Before using a burette, it is necessary to clean the burette by rinsing with water. When filling a burette it is advisable to take it out of the burette clamp and hold it at a comfortable level.

Do NOT climb on a stool to fill a burette.

Do NOT use a large filter funnel to fill a burette.

Rinse the burette well with **water**. Close the tap. Add about 5–10 mL of water. Hold the burette in the horizontal position. Shake and rotate the burette so that the surface is wetted. Open the tap and drain.

Repeat the procedure but **rinse the burette** with the **titrant**. Use no more than 5 mL of titrant. Perform this rinsing procedure twice.

Close the tap and fill with the titrant making sure the level of solution is well above the zero mark.

Open the tap and run out some of the titrant quickly. This will usually **remove all air bubbles in the tube below the tap**. When there are no more air bubbles, close the tap. If there are air bubbles which won't move, consult your demonstrator.

Slowly open the tap and drain so that the bottom of the meniscus is at the desired graduation mark. This does not necessarily have to be the zero mark. Note this reading to ± 0.05 mL, and then begin the titration.

Remember to avoid parallax error.

In carrying out a titration, hold the tap with the left hand for a right-handed person; and the right hand for a left-handed person. The tap is held by the thumb, index and middle fingers; positioned from behind the tap.

Ask your demonstrator to show you the finer details of titration, e.g., how to ‘cut’ a drop.



Materials Needed

~0.1 M NaOH (different solutions for different workspaces)
standardized 0.1 M HCl solution
phenolphthalein

50.0 mL burette
250 mL conical (Erlenmeyer) flasks
25.00 mL pipette
pipette filler

100mL beaker
250 mL beaker
500 mL beaker



Do not waste the standardised acid and base solutions. These solutions take considerable time and effort to prepare. Take no more than the volume listed in the instructions.

Method

Titrate the sodium hydroxide solution with the HCl standardised solution

1. Collect no more than 100 mL of the appropriate NaOH(aq) provided for your workspace into a **clean, dry, labelled** 250 mL beaker.
2. Rinse the burette with 10 mL of water followed by about 10 mL of the sodium hydroxide solution, discarding the waste solution into a 500 mL beaker. Rinse again with 10 mL of the sodium hydroxide and fill the burette with the sodium hydroxide solution to between 1–5 mL from the zero mark at the top. Remember to **take the burette out of the clamp and hold it in front of you** when you fill it. If you have used a small plastic funnel, **remove the funnel now**. Record the initial reading to the nearest 0.05 mL on your results page.

- Collect no more than 100 mL of the HCl(aq) in a **clean, dry, labelled** 150 mL beaker. Record the accurate concentration on the results sheet in pen.
- Pipette 25.00 mL of the HCl(aq) into a 250 mL conical flask and add 1 or 2 drops of phenolphthalein solution to the flask.
- Slowly add the sodium hydroxide solution to the HCl, swirling as you do so. You should see a pink colour where the NaOH enters the acid, which fades as you swirl the solution.

You must **rinse the inside of the flask** with water from the wash bottle after the first appearance of the pink colour. This washes all splashes off the side of the flask, and will give a more accurate end point. Ask your demonstrator to show you this technique if you are unsure. Swirl the solution in the flask but do not shake it, to avoid a significant amount of atmospheric carbon dioxide dissolving and affecting the endpoint.

- Continue titrating until the first permanent **pale** pink colour is obtained. The pink colour should persist for a period of at least 30 seconds in order to be regarded as permanent. Record in pen on the results page the final reading to the nearest 0.05 mL of sodium hydroxide required to reach the end point. Show the end point to your demonstrator and have your titre (titration volume) checked before you proceed.

You can use the **wash bottle to cut a drop** in the subsequent titrations for a more accurate end point. Ask your demonstrator to show you this technique.

- Repeat the titration on fresh samples of HCl(aq) until at least **three** concordant results are obtained. Titres should not differ from the average by more than 0.1 mL to be considered concordant.
- Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.
- Calculate the concentration of the sodium hydroxide solution.

WRITE UP

If you are unable to complete the write up during the laboratory session, you should:

- have your demonstrator sight the work you have completed and initial your laboratory manual
- complete the calculations and your work will be checked by your demonstrator during your next lab class.

RESULTS – Standardisation of a Base

Your name: _____ Lab day/time: _____ Workspace: _____

TITRATION: HCl with NaOH

Which NaOH solution did you use? (sample A, B, or C?) _____

Accurate concentration of hydrochloric acid used _____

Titration	Initial Reading (± 0.05 mL)	Final Reading (± 0.05 mL)	Titre (± 0.1 mL)	Used in calculating average? (Yes/No)
1				
2				
3				
4 (if needed)				
5 (if needed)				
				Demonstrator's Initials

CALCULATION – Concentration of sodium hydroxide solution

Note: Three concordant titres (*i.e.* all within ± 0.1 mL of their average) should be obtained. Show your calculations for the amount of each reactant in the space provided below and hence determine the concentration of sodium hydroxide.

Average of concordant titres = _____ mL

Chemical equation for reaction in the titration:

Calculations:

Concentration of NaOH = _____ mol L⁻¹ (with appropriate number of significant digits!)

[For CHEM1031 and CHEM1051 only]

Describe how your calculations would have changed if sulfuric acid was used instead of hydrochloric acid.

Feedback (Demonstrator to complete – Skills assessment – tick ALL that apply)**Core Skills**Analytical glassware : pipette – correct useTitration: setup (rinse, fill, no funnel or bubble) swirling during addition & pale endpoint**Non-core Skills**[CHEM1031/51 only] Applying chemical principles: correct logic and explanation used appropriate principle(s)Chemical calculations: correct arithmetic correct method correct sig. fig. on resultsExperimental accuracy: concentration within 5% of expected value[CHEM1031/51 only] Mastery (Applying chemical principles): clear & concise explanation(s)Mastery (Experimental accuracy): concentration within 1% of expected valueProfessionalism: [CHEM1031/51 only] answers and discussion in pen locker and contents left cleanRecording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manualSafety awareness: safety eyewear on at all timesTime management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

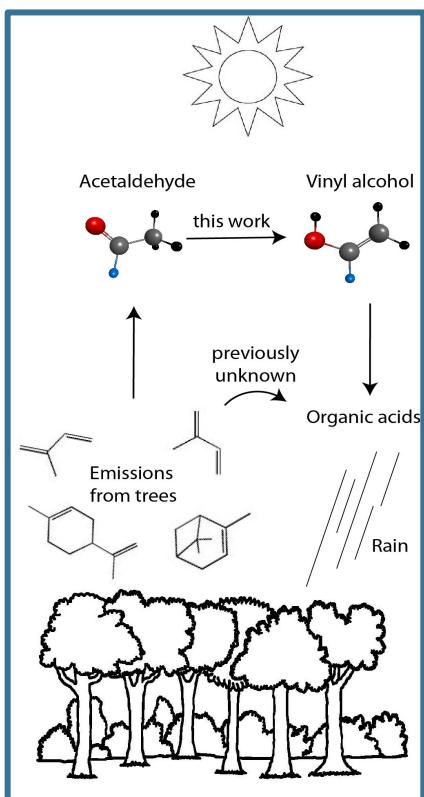
Cool Research in the School of Chemistry

Professor Scott Kable

and his group are using lasers to study
combustion, the atmosphere, and space.

The research in my group is focused on discovering new intermediates and uncovering new mechanisms in complex chemical systems. We have discovered new intermediates in the atmosphere, surprisingly selective radical reactions in combustion, and characterized a completely new photochemical mechanism.

Projects can be tailored to be experimental and/or computational. Experiments utilise multiple lasers, high vacuum and imaging technology and timescales from femtoseconds to minutes. Theory employs the most modern methods on our own machines, or the Australian supercomputer.



Atmospheric chemistry

100 M tonnes of organic acids are found in the atmosphere but their source is unknown. We recently discovered that acetaldehydes photo-tautomerize in the atmosphere. The enol form reacts rapidly to form organic acids. Inclusion of this pathway in atmospheric models increase the rate of acid formation by **10 times!** (see picture above)

Radicals in chemistry

Free radicals are the scavengers of the atmosphere, reacting with all organic pollutants. They are chain propagators in all combustion systems and are found throughout space. But they are very difficult to isolate and study. We have developed techniques for synthesizing and probing radicals leading to the discovery of new radicals and uncovering surprisingly selective reactions. This selectivity will change existing mechanisms of combustion and the atmosphere.

Photochemistry mechanisms

“Roaming” refers to a new mechanisms where two photofragments orbit each other and re-react to produce completely unexpected products. We discovered the second ever roaming mechanism in 2006; since then roaming has been found in combustion, the atmosphere and is now believed to be a ubiquitous pathway in photochemistry.

The Laser Spectroscopy Group

The Laser Spectroscopy Group is a collaborative group led by Profs Schmidt and Kable. The group moved from U. Sydney at the start of 2014 into new laboratories on Level 4 of the Chemical Sciences tower.

If you are interested in a 3rd year, Honours or Ph. D. project in this area drop me an email (s.kable@unsw.edu.au), tour the lab, and talk to our students and post-docs.

VOLUMETRIC ANALYSIS – BACK TITRATION

Aim

To determine the molar mass of an unknown inorganic carbonate and hence deduce the identity of the compound from a list of possible substances.

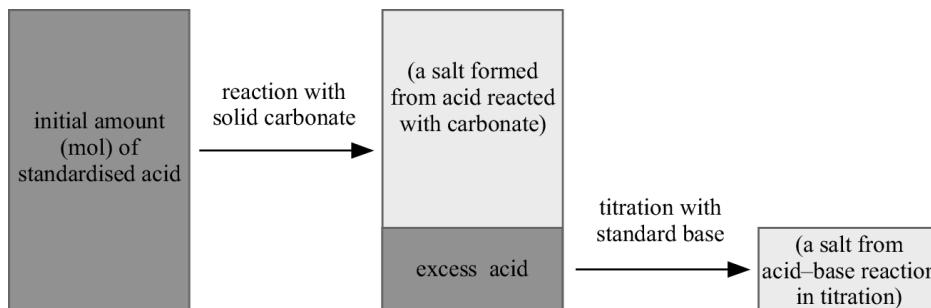
See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

Usually the aim of volumetric analysis is to determine the concentration of one solution by the addition of a stoichiometric amount of a second solution from a burette (a titration). If the volumes of each of the two solutions are accurately known, and the concentration of one is also accurately known, the concentration of the other solution can be determined. In a previous experiment you have carried out this sort of titration. This experiment introduces two variations on this method.

In the previous titration experiment the reaction in the titration was between a strong acid and a strong base. This reaction happens rapidly, and completely, in the flask as one solution is added from the burette. To get an accurate result in any titration you need a fast and complete reaction as the titration proceeds. Reactions involving solids which dissolve slowly, or reactions producing gases are usually not suitable for use in a titration. There is a way to get around this limitation by letting the slow reaction complete using a known amount of reactant in excess of what is needed, and then doing a simple titration to determine how much of the reagent is left over. This is called a **back titration**.

For example, say you wanted to determine the amount of copper(II) carbonate in a rock. Carbonates react with acid, but the reaction generates $\text{CO}_2(\text{g})$ which makes seeing the end-point difficult, and the rock will contain other insoluble substances which will also interfere with seeing the end-point. To avoid these problems, the rock is reacted with a known volume of a standardised acid, so that all the CuCO_3 reacts and some acid is left over (the carbonate is the limiting reactant, and the acid is in excess). This reaction can be carried out slowly, with heating to remove CO_2 , and the final solution might be filtered to remove insoluble material. The final solution can be titrated with standardised base to determine the concentration of acid in it. You now know the amount of acid added to the carbonate (from its concentration and volume) and the amount of acid unreacted or leftover (from its volume and concentration, determined by the titration). The difference between these two values gives the amount of acid which reacted with the carbonate, which allows you to determine the amount of copper(II) carbonate in the rock.



This experiment is very similar to the example just described. You will react a weighed amount of a metal carbonate with excess standardised acid, allow that reaction to complete, then titrate the excess acid with standardised base. However in this experiment you will not know what the metal in the metal carbonate is and you will be asked to determine the molar mass of the metal carbonate. You do this by calculating the amount (mol) of carbonate which reacted with the acid, and since amount = mass/molar mass, from the weighed mass of carbonate you can calculate its molar mass.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
0.1 M NaOH		
1 M HCl		
unknown carbonate		As you do not know the identity of the sample, treat it as potentially hazardous. Substances which need special precautions are labelled as "HARMFUL" or "IRRITANT"
phenolphthalein		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Techniques

Review the techniques of weighing by difference, using a pipette and performing a titration, and watch the videos online (links on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these techniques.

Materials Needed

0.1 M standardised NaOH	50.00 mL burette	250 mL conical (Erlenmeyer) flasks
1.0 M standardised HCl	25 mL pipette	250 mL beaker
unknown carbonate (solid)	10 mL pipette	150 mL beaker
phenolphthalein	100 mL volumetric flask	analytical balance steam bath



Do not waste the standardised acid and base solutions. These solutions take considerable time and effort to prepare. Take no more than the volume listed in the instructions.

Method

1. Weigh by difference between 0.45 – 0.55 g of the carbonate assigned to your workspace into a clean 250 mL conical flask (clean, but may be wet with water). Record the mass you weigh out and the identifier of the carbonate.
2. Obtain no more than 40 mL of standardised 1.0 M HCl in a **clean, dry, labelled** 150 mL beaker.
3. Use a 25 mL pipette to transfer 25.0 mL of the 1.0 M HCl into the flask containing the solid carbonate. Allow the reaction to complete, then warm the flask on a steam bath for five minutes to expel dissolved CO₂.
4. After the solution has cooled, transfer it quantitatively into a clean 100 mL volumetric flask. Make up to the mark with tap water and invert the flask to thoroughly mix the contents.
5. Obtain a 50 mL burette and no more than 100 mL of standardised 0.1 M NaOH (in a **clean, dry, labelled** 250 mL beaker).
6. Rinse the burette with a few mL of the NaOH (take care not to spill or splash the NaOH) and then fill the burette with the NaOH.
7. Use a 10 mL pipette to transfer 10.0 mL of the solution in the volumetric flask into a clean 250 mL conical flask (does not have to be dry).
8. Add two drops of phenolphthalein to the solution and titrate with the NaOH. Record the endpoint in the results section.
9. Rinse the conical flask with water and repeat steps 7 and 8 until concordant results are obtained (3 results in a 0.2 mL range).
10. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.

Blank page.
(Can be used for working or calculations)

School of Chemistry, UNSW

Coversheet for Submission of Individual Report

Report: Volumetric Analysis – Back Titration

Your course: CHEM1011 CHEM1031 CHEM1051

Student Name (full name): _____

Student Number (e.g., z1234567): _____ Lab. class: day/time: _____

Demonstrator: _____ Workspace: _____

Date and Time Submitted: _____

In preparing this assessment task I have followed the Student Code of Conduct. I certify that I have read and understand the University requirements in respect of student academic misconduct outlined in the Student Code of Conduct and Annexure 1 of the Student Misconduct Procedures. I declare that this assessment item is my own work, except where acknowledged, and has not been submitted for academic credit previously in whole or in part. I acknowledge that the assessor of this item may, for assessment purposes: (1) provide a copy to another staff member of the University; (2) communicate a copy of this assessment item to a plagiarism checking service (such as Turnitin) which may then retain a copy of the assessment item on its database for the purpose of future plagiarism checking. I have retained a copy of this, my assignment, which I can provide if necessary. By signing this declaration I am agreeing to the statements and conditions above.

Student signature and date: _____

Full UNSW policy concerning student misconduct is available at: <<https://student.unsw.edu.au/conduct>>

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Core Skills

Analytical glassware: pipette – correct use volumetric flask - correct use (last chance)

Heating materials: steam bath – correct use

Titration: setup (rinse, fill, no funnel or bubble) swirling during addition & pale endpoint

Weighing: analytical balance - weighing by difference (last chance)

Non-core Skills

Applying chemical principles: correct logic and explanation - [CHEM1031/51 only] used appropriate principle(s)

Chemical calculations: correct arithmetic correct method correct sig. fig. on results correct units

Chemical equations: balanced - atoms & charge correct formulae correct states of matter no spectator ions

Experimental accuracy: acceptable molar mass

Mastery (Experimental accuracy): accurate molar mass

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

RESULTS

Unknown carbonate identifier _____

Mass of carbonate/g _____

Standard HCl concentration/M _____

Standard NaOH concentration/M _____

TITRATIONS

Titration	Initial Reading /mL (± 0.05 mL)	Final Reading /mL (± 0.05 mL)	Titre /mL (± 0.1 mL)	Used in calculating average? (Yes/No)
1				
2				
3				
4 (if needed)				
5 (if needed)				
6 (if needed)				
Average titre and uncertainty				

Space for calculations of the molar mass is provided on the next page.

CALCULATIONS

Chemical reactions

Reaction of $\text{H}^+(\text{aq})$ with $\text{CO}_3^{2-}(\text{aq})$ (step 3 of the method)

Reaction of $\text{H}^+(\text{aq})$ with $\text{OH}^-(\text{aq})$ (step 8 of the method)

Hints for doing the calculations

First calculate the amount (mol) of acid transferred in step 3. This is the initial amount of acid before some of it reacts with the carbonate ions. Use the volume and concentration of base in the titration (steps 7 and 8) to calculate the concentration and then the amount (mol) of acid in the 100 mL volumetric flask. The difference between these two amounts of acid is the amount of H^+ which reacted with the carbonate. Use the appropriate chemical reaction to calculate the amount (mol) of carbonate in the weighed sample. Since you know that amount (mol) = mass/molar mass, from the weighed mass of carbonate you can calculate its molar mass and deduce its identity using the adjacent table.

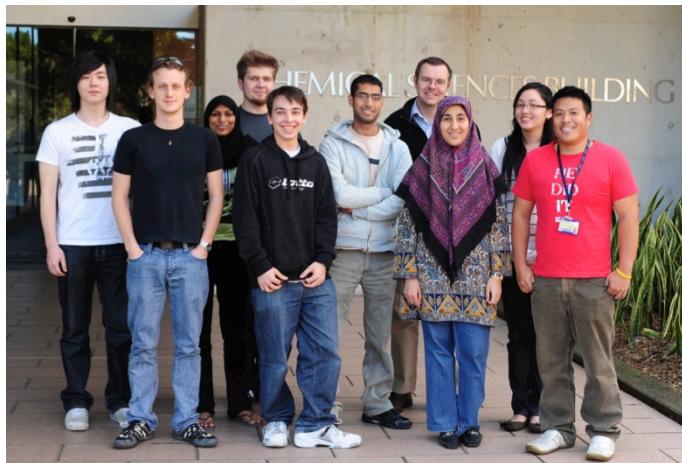
Compound	Molar Mass / g/mol
Li_2CO_3	73.9
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	286.1
K_2CO_3	138.0
MgCO_3	84.3
CaCO_3	100.0
SrCO_3	147.6
BaCO_3	197.3

Show your calculations here.

Measured molar mass (and units)		Identity of carbonate	
------------------------------------	--	-----------------------	--

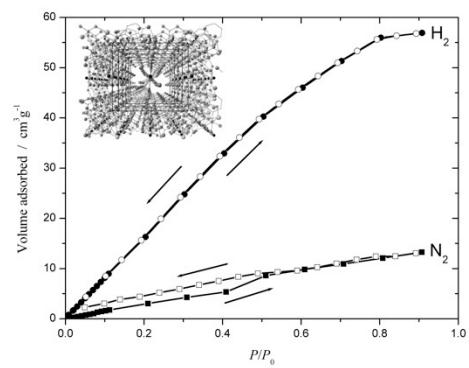
[For CHEM1031 and CHEM1051 only]

Why was copper(II) carbonate not selected as an 'unknown' compound for this experiment?

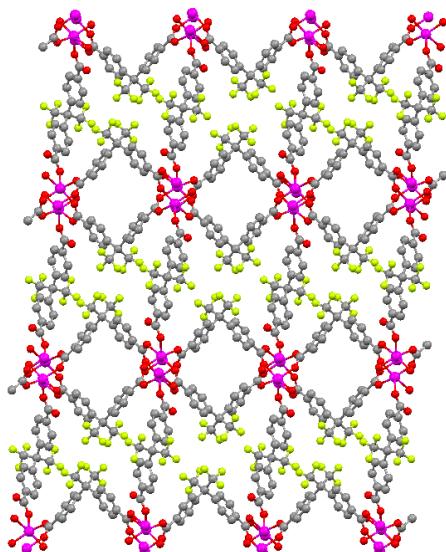


Cool Research in the School of Chemistry

Assoc. Professor John Stride
and his group are studying
Molecular Devices:
Functional Materials



A Cu-based MOF displaying a 5-fold preference for H₂ over N₂ due to size-exclusion in the very small pore sizes in the



One of the MOFs that displays large pores by virtue of the phase separation of the fluorinated and hydrocarbon regions of the ligands.

Highly porous materials called metal-organic frameworks (MOFs) are sponge-like materials with pores of molecular dimensions. They have the ability to trap guest molecules including hydrogen (for use as a fuel) and carbon dioxide and other polluting and greenhouse gases. The Stride group has developed a large number of new MOFs. The adsorption and magnetic characteristics of these materials are of particular interest with materials that are both highly selective and able to convey occupancy via electronic correlations within the host being sought after. To date they have obtained materials that have shown either high selectivity or strong magnetic coupling and their efforts are now focussed on cross-fertilising these properties.

It is hoped that these MOFs will play two roles in solving the world's climate problems: by capturing emissions in sequestration technologies and by solving the problem of effectively storing low-polluting hydrogen which is a highly explosive gas at ambient temperature and pressure. With the levels of atmospheric carbon dioxide set to rise from 390 ppm (by volume) to 500 ppm within 40 years, therefore leading to rising sea-levels, a solution cannot come fast enough.

For more information about this research see:
<http://www.chemistry.unsw.edu.au/research/research-groups/stride-group>

CORROSION

[In semester 1, 2018 only CHEM1031 and CHEM1051 will do this experiment.]

Aim

The aim of this experiment is to investigate the corrosion of iron (as iron nails) on its own and when in contact with other metals such as Cu, Sn, Zn and Al. You will determine which metals enhance the corrosion of iron and which inhibit its corrosion. In the second part of the experiment, measurements will be made of the electric current flow during a corrosion (redox) process.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

Corrosion is the conversion of a material into some other chemical form by reaction with chemicals in the environment. The most common form of corrosion is the conversion of a structurally strong metal into a brittle ionic compound. If the metal was part of a building, or bridge, or ship, corrosion can have catastrophic consequences. Corrosion also results in the conversion of an electrically conducting metal into a non-conducting ionic solid, which again can have significant practical consequences. Metal compounds released by corrosion can be toxic or (in the food industry) accelerate spoilage of foods. Corrosion prevention is one of the important problems confronting science. This experiment will concentrate on the corrosion of iron as it is the most widely used metal in construction and corrosion is a major problem for iron.

The conversion of a metal into an ionic compound involves the loss of electrons from metal atoms to form positive ions (cations), e.g., $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$. This is an example of an **oxidation** (electron loss) process. Every oxidation reaction must be accompanied by an electron gain reaction (a **reduction** reaction), and in most corrosion situations the reduction involves oxygen and water: $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$. The combination of a reduction reaction with an oxidation is called a **redox reaction** (e.g., $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 2\text{Fe(s)} \rightarrow 4\text{OH}^-(\text{aq}) + 2\text{Fe}^{2+}(\text{aq})$). You will learn more about redox reactions later in Chemistry A, but this experiment does not rely on more knowledge than what is provided in this introduction.

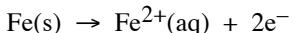
Some metals are much easier to oxidize (e.g., zinc, aluminium, magnesium), while others are difficult to oxidize and so do not corrode (e.g., copper, silver, gold). These less reactive metals are often rarer in nature and thus more expensive.

Paints and other surface coatings can be used to keep oxygen and water from reaching the metal surface but there are other methods which use the chemical properties of the metals themselves. Some metals (e.g. aluminium, zinc) are much less reactive than would be expected, owing to the presence of a strongly adherent protective oxide layer. **Galvanizing** is the coating of iron with a thin layer of zinc. The zinc forms a tough oxide layer which stops oxygen and water reaching the iron. Small scratches in the zinc layer result in a new oxide layer forming because zinc is more easily oxidized than iron and corrosion is halted. The converse of this is seen in 'tin' cans where iron is coated with tin which is less easily oxidized than the iron and also forms a protective oxide layer. If the tin coating is scratched, the underlying iron corrodes rapidly. In this experiment you will see the corrosion process for iron in contact with zinc, and iron in contact with tin.

By immobilizing an iron nail, in contact with a second metal, in agar gel in a test tube, it is possible to locate the sites of the oxidation and reduction reactions. In each of the reactions studied, the reduction involves O_2 and H_2O :



The oxidation which occurs will convert a metal into a cation. The oxidation **may** involve the iron:



or the metal in contact with the iron. Which metal actually oxidizes will depend on the relative reactivity of the Fe and the metal concerned. Based on your observations you should be able to determine which metal is being oxidized in each of the test tubes.

Since the agar gel contains phenolphthalein which turns pink in the presence of OH^- , any pink colouration in the gel indicates a site at which reduction occurs (forming OH^-). In all cases, since the reduction is the same reaction, a pink colour will be observed somewhere in the test tube if corrosion is occurring. **A PINK colour means reduction of O_2 is occurring.** If reduction is happening, oxidation must also be happening.

The gel also contains potassium ferricyanide (potassium hexacyanoferrate(III), $K_3[Fe(CN)_6]$), which turns blue on reaction with Fe^{2+} :



Thus a **BLUE colour indicates a site at which Fe is being oxidized**. Since the Fe is not oxidized in all cases, the blue colour will not appear in all tubes. However if you see a pink colour in a tube you know reduction is occurring, and therefore something must be being oxidized, and if it is not the iron, then it must be the other metal.

In the second part of this experiment you will confirm which metal is oxidized by determining the direction of electron flow between iron and two other metals by determining the direction of electron flow using a voltmeter.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

The web page where you get your safety information has a link to VIDEOS about several of the techniques you will use in this experiment. **You should watch these videos before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
agar		
phenolphthalein		
copper foil		
zinc foil		
aluminium foil		
tin foil		
magnesium ribbon		
iron nails		
0.1 M $K_3[Fe(CN)_6]$		
NaCl		
acetone		

2 M H ₂ SO ₄		
3% NaCl		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

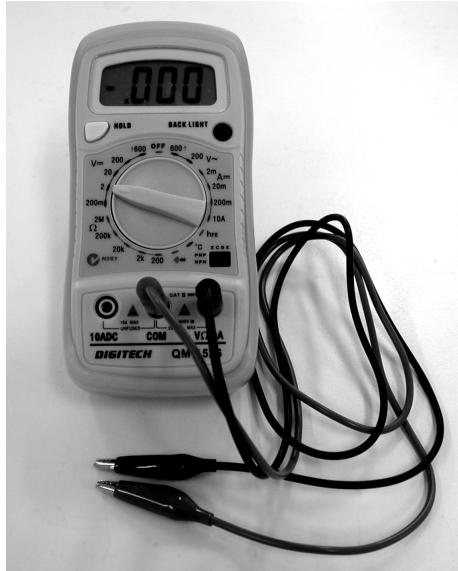
Demonstrator's signature and date	
-----------------------------------	--

Techniques

Read the instructions below and watch the videos online (links on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these techniques.

TECHNIQUE – USING A DIGITAL MULTIMETER

- (i) Switch the meter to the 2 V DC range.
- (ii) Attach the lead from the VΩA terminal to one electrode of the cell and the other lead from the COM terminal to the other metal in the cell. Make sure there is good contact between the clips and the metal electrodes – squeeze and twist the clips to get a steady reading on the meter.
- (iii) The voltage can be read directly from the digital display.
- (iv) **A positive reading on the meter indicates that the electrode connected to the VΩA terminal is positive with respect to the electrode connected to the COM terminal.**



Materials Needed

watch glasses	steel wool	copper foil	NaCl
10 mL measuring cylinders	cotton wool	zinc foil	acetone
250 mL beaker	agar	aluminium foil	2 M H ₂ SO ₄
bunsen burner, wire gauze, heat-proof mat	phenolphthalein	tin foil	3% NaCl
top loading balances	0.1 M K ₃ [Fe(CN) ₆]	degreased iron nails	hammers
five 15 x 1.5 cm (narrow) test tubes	magnesium ribbon	iron strips	



Make sure all the glassware you use is clean.

This experiment can fail if the glassware used to prepare and hold the solutions and gels is not clean. Make sure you rinse all glassware thoroughly under running water in a sink before use. If you use detergent, make sure you rinse away ALL the detergent (some detergents are basic). The glassware does not have to be dry.

Method – Part A



Avoid burns when using Bunsen burners.

The vessel you are heating will be hot as will the metal gauze. Remove the vessel with an insulating material such as cloth or paper to protect your fingers. Make sure the gauze has cooled before removing it from the rod.



Check rubber tubing on Bunsen burners before use.

Check the whole rubber tube for splits or cracks and take it to the issue room if it is damaged in this way. Do not use tubing which is split or cracked.

Cleaning the Iron Nails

1. Place 5 iron nails into one test-tube containing 15 mL of 2 M H_2SO_4 for about 2 minutes.
2. Boil about 50 mL of water in a 250 mL beaker, decant (pour off) the acid from the nails in the previous step, rinse them several times with tap water and then carefully add them to the boiling water. After 1 minute (no longer!) turn off the Bunsen. Remove the nails one at a time as required, using a pair of metal tongs.

Preparation of Agar Gel

3. On separate watch glasses weigh 0.8 g of agar and 5.0 g of NaCl using top loading balances. Note: the mass of agar required may vary depending on the supplier of the agar. Check with your demonstrator that 0.8 g is the mass required for your particular class.
4. Thoroughly rinse a 10 mL measuring cylinder and two tall, narrow test tubes. Use the measuring cylinder to dispense 2 mL of phenolphthalein solution into one test tube and 1 mL 0.1 M potassium hexacyanoferrate(III) $[\text{K}_3\text{Fe}(\text{CN})_6]$ into the other test tube, ready for step 9 below.
5. Wash a 250 mL beaker with a little detergent and rinse well before use.
6. Heat to boiling about 100 mL of demineralized water in the cleaned 250 mL beaker.
7. Move the Bunsen aside, add the agar to the water, and then resume heating with stirring until the agar has completely dissolved.
8. Turn off the Bunsen and add the NaCl to the hot solution and stir until it dissolves.
9. Add 2 mL phenolphthalein indicator and 1 mL 0.1 M potassium hexacyanoferrate(III) $[\text{K}_3\text{Fe}(\text{CN})_6]$ (from step 4), **stir the solution well**. Proceed with the following steps while the mixture is cooling. The mixture should be yellow, not pink, green, blue, or colourless. Consult your demonstrator if the mixture is not yellow.

Note: The solution of $\text{K}_3[\text{Fe}(\text{CN})_6]$ and phenolphthalein is known as 'ferroxyl' indicator and will give a blue colour with Fe^{2+} (ferrous) ions and a pink colour with OH^- (hydroxyl) ions.

Treatment of Cleaned Nails

10. **Thoroughly rinse** five test tubes (tall narrow ones). Label them 1–5. In tube 1 place one of the cleaned nails. In tubes 2–5 you will place a cleaned nail wrapped in a piece of metal (as described below). It is important that the piece of metal fits tightly around the nail.
11. Using a fresh nail (NOT ONE CLEANED IN THE STEPS ABOVE) and a hammer, punch a hole in a piece of tin foil and then insert one of the **cleaned** nails through the hole, making sure that the contact between the nail and tin is good.
12. In a similar manner, punch holes in pieces of zinc foil, aluminium foil and copper foil and insert a cleaned nail into the holes. Place each nail+metal into the appropriate labelled test tube (tubes 2–5). Pour sufficient indicator gel (which should now be just warm to the touch) into each of the test tubes (tubes 1–5) to cover the nail, avoiding air bubbles. If the indicator gel has cooled until it is too viscous to pour, warm it on a steam bath until it can be poured. Look at the picture in the results section if you are not certain about what is required.
13. Place the test tubes 1–5 in a rack and examine them throughout the remainder of the laboratory session but do not shake or move them around.

14. Observe any coloured regions which develop in the gel and indicate these (using colored pencils or pens) on the diagram on the results page. Carry out part B while you are waiting for the reactions to proceed.
15. Discard the agar mixture and nails into the bucket provided in the fume cupboard. You may need to warm the test tube in a steam bath to soften the gel.



**DO NOT EMPTY THE GEL AND NAILS INTO THE SINKS!
PUT IT IN THE BUCKET IN THE FUMECUPBOARD.**

Method – Part B

NOTE: There are details about using a multimeter in the technique section on page 43. If you have never used a multimeter before you should look over that material now.

1. Clean the iron, copper and magnesium strips with steel wool, then rub them with a piece of cotton wool soaked in acetone to remove any grease.
2. Stand one of the iron and one of the copper strips in a 100 mL beaker. By means of alligator clips connect the multimeter between the two metals. Add 30 mL of 3% sodium chloride solution to the beaker, making sure that the two metals do not touch and that the alligator clips are not wet. Set the multimeter to the 2V DC range and determine which of the metals is positive. A positive reading on the multimeter indicates that the metal connected to the VΩA terminal is positive with respect to the electrode connected to the COM terminal. Determine the direction of electron flow. Hint: electrons flow from the external circuit into the metal which is positive compared to the other metal.
3. Connect a piece of magnesium in place of the copper electrode. Determine which is the positive electrode and the direction of electron flow.
4. Rinse the strips of metal and return them to where you got them. **DO NOT THROW OUT THE METAL STRIPS.**
5. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.
6. Complete the table and answer the questions in the results section.

RESULTS – Corrosion

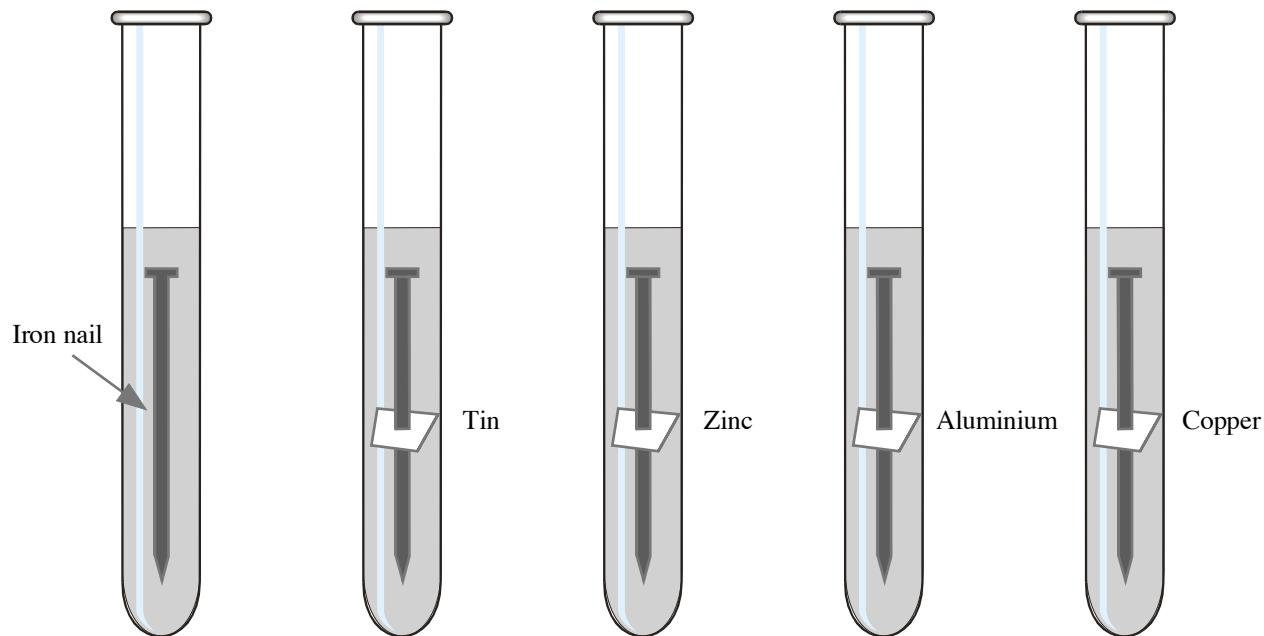
Your name: _____

Lab day/time: _____

Workspace: _____

Part A

On the diagram below sketch the regions where colour has developed and what colour each region is for each test tube. You MUST do this in the laboratory in order to be able to complete this report.



Each test tube contains agar gel with phenolphthalein and potassium ferricyanide

Use your observations to complete the following table. The first line has been completed to show you what is required.

Metal(s)	Presence and location of pink colour	Presence and location of blue colour	Ionic equations	
			Oxidation reaction	Reduction reaction
Fe	along the shaft	tip and head	$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$	$\text{O}_2(\text{aq}) + 2 \text{H}_2\text{O(l)} + 4\text{e}^- \rightarrow 4 \text{OH}^-(\text{aq})$
Fe/Sn				
Fe/Zn				
Fe/Al				
Fe/Cu				

Based on your observations, which of the metals (Sn, Zn, Al, Cu) inhibited the corrosion of the iron?.

Galvanised iron is iron or steel coated with zinc. From your observations, explain how galvanising iron will reduce corrosion, even if the zinc surface coating is damaged.

Based on your observations, would it be a good idea to use an iron bracket to hold a copper pipe to a wall? Explain what would happen.

Part B

Metal Pair	Which metal is positive?	Direction of electron flow through the meter	Metal which is oxidized?
Fe + Cu		From to	<input type="checkbox"/> Fe <input type="checkbox"/> Cu
Fe + Mg		From to	<input type="checkbox"/> Fe <input type="checkbox"/> Mg

Are your observations for Fe + Cu in part B consistent with what you observed for these metals in part A? Briefly justify your answer.

Corrosion accelerated by biochemical processes [CHEM1051 only]

There are species of bacteria which can convert sulfur into H_2S , and other bacteria which can convert H_2S (via a reaction with O_2) into sulfuric acid. Pipes made of steel can corrode rapidly via reaction with the sulfuric acid.

What common biochemicals could serve as the original source of the sulfur?

What urban pipe network is particularly at risk from this form of corrosion?

Would pipes made of concrete (which contains calcium hydroxide) be damaged under these conditions?

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)**Core Skills**

Heating materials: steam bath – correct use

Non-core Skills

Applying chemical principles: correct logic and explanation used appropriate principle(s)

Chemical equations: balanced - atoms & charge correct formulae correct states of matter no spectator ions

Describing chemical changes: correct use of terms, e.g., precipitate, supernatant

Mastery (Applying chemical principles): clear & concise explanation(s)

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

INTERMOLECULAR FORCES

Aim

The aim of this experiment is to investigate aspects of bond types and intermolecular forces both experimentally (by comparison of solubility in a variety of solvents and the effect of added solutes on liquids) and by simulation (through a study of vapor pressure).

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

Pure substances

Pure substances may be ionic, metallic, molecular or network covalent. This experiment is concerned with ionic and molecular substances. Ionic substances are held together by electrostatic attraction between oppositely charged ions. In molecular substances, there are three major intermolecular forces which may be present:

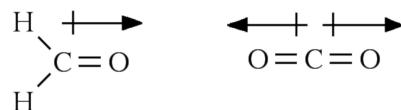
- dispersion forces (also known as induced dipole or London forces)
- dipole/dipole interactions
- hydrogen bonding.

Ionic bonds are usually stronger than any of these intermolecular forces.

Dispersion forces are present in all molecules and arise due to temporary distortion of the electron clouds of the molecules when two molecules approach one another. The larger the molecule (and hence the larger its polarizability), the greater the distortion and the stronger the dispersion force. So dispersion forces are larger between molecules containing larger numbers of atoms, or molecules containing larger atoms, e.g., the dispersion forces between F_2 molecules are smaller than between I_2 molecules (remember that atomic size increases down a group in the periodic table and F and I are both in the halogen group, but iodine is near the bottom while fluorine is at the top of the group).

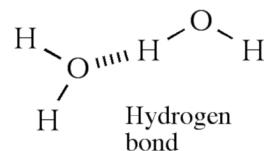
Dipole/dipole attraction occurs in molecules which have permanent dipoles due to polar bonds within the molecule and a lack of symmetry within the molecule. For example, H_2CO is polar but CO_2 is not, even though both contain polar C=O bonds.

Since CO_2 is linear, the individual bond dipole vectors add up to zero, making the molecule non-polar. In H_2CO , there is an overall residual dipole, making the molecule polar. Remember that dipole moments (of bonds and molecules) are vector quantities – they have size and direction.



In molecules of comparable size, one polar and the other non-polar, the intermolecular forces in the polar molecule would be expected to be greater than in the non-polar molecule, since there is a permanent attraction between molecules (which supplements the dispersion forces present between all molecules, polar or non-polar).

For **hydrogen bonding** to occur, a molecule needs a hydrogen atom bonded to a highly electronegative atom such as oxygen, nitrogen or fluorine, leading to a highly polar bond. When combined with the small size of the hydrogen atom, close approach between the hydrogen in one molecule and the electronegative atom in the other leads to a strong intermolecular attraction.



In molecules of comparable size, one with hydrogen bonding will in general have stronger intermolecular forces compared to a molecule which cannot form hydrogen bonds.

The strength of the intermolecular forces between molecules is reflected in the boiling point and vapor pressure of a substance. The stronger the attractive force, the more energy is required to separate the molecules from one another and the higher the boiling point and the lower the vapor pressure at a given temperature. Remember when a substance boils or vaporizes, it is the intermolecular forces which are overcome – it is **NOT** the covalent bonds which are breaking. The covalent bonds within the molecules remain intact as the substance changes its physical state.

In the simulation part of this experiment, the vapor pressures of a number of covalent molecules will be compared and inferences drawn regarding the relative strengths of their intermolecular forces. The sets of compounds which you will compare are:

methanol <chem>CH3OH</chem>	ethanol <chem>CH3CH2OH</chem>	propanol <chem>CH3CH2CH2OH</chem>
--------------------------------	----------------------------------	--------------------------------------

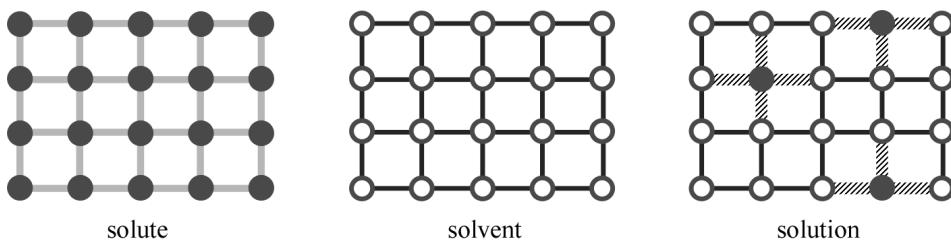
2-chloropropane $\begin{array}{c} \text{Cl} \\ \\ \text{H}_3\text{C} - \text{C} - \text{CH}_3 \\ \\ \text{H} \end{array}$	propan-2-ol or 2-propanol $\begin{array}{c} \text{OH} \\ \\ \text{H}_3\text{C} - \text{C} - \text{CH}_3 \\ \\ \text{H} \end{array}$	propanone or acetone $\begin{array}{c} \text{O} \\ \parallel \\ \text{H}_3\text{C} - \text{C} - \text{CH}_3 \end{array}$
--	--	---

pentane <chem>(CH3CH2CH2CH2CH3)</chem>	hexane <chem>(CH3CH2CH2CH2CH2CH3)</chem>	heptane <chem>(CH3CH2CH2CH2CH2CH2CH3)</chem>
---	---	---

Solutions (mixtures)

A solution is a mixture of one or more solutes in a solvent. If solute dissolves sufficiently to produce a solution with a concentration above 0.1 mol/L it is said to be **soluble** in that solvent. If two liquids are mutually soluble they are said to be **miscible**.

As a solute dissolves in a solvent the forces between solute particles are replaced by solute–solvent forces, as are some of the forces between the solvent particles. Whether or not a solute dissolves in a particular solvent is largely determined by the energy change in that process and this depends on the relative strengths of the solute–solvent forces compared to the solute–solute and solvent–solvent forces.



In the diagram immediately above (showing the different types of interaction between the particles, not their actual positions) the new forces are shown as shaded thick lines in the solution picture. To form a solution from the separated solute and solvent:

1. requires energy to overcome the solute – solute bonds or forces
2. requires energy to overcome the solvent – solvent attractions
3. produces energy from the formation of solute – solvent attractions.

If the total energy required in steps 1 and 2 is much larger than the energy released in step 3 then it is not likely that the solute will dissolve in the solvent.

In forming aqueous solutions the energy required to separate the hydrogen bonded water molecules is large and a solute will dissolve only if water molecules will be strongly attracted to it. Substances which dissolve in water are usually either:

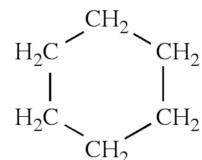
- ionic – they dissociate to form ions and the resulting ion – dipole (ion–water) forces are strong, or
- molecular substances which dissociate to form ions, or
- molecules which can form hydrogen bonds – and thus form hydrogen bonds to the water molecules, or
- very polar molecules – and thus are attracted to water molecules by dipole – dipole forces.

Substances composed of non-polar or slightly polar molecules are usually insoluble in water because they are not attracted strongly enough to water molecules to make up the energy needed to separate the water molecules when the solute dissolves.

Substances composed of low polarity molecules will dissolve in non-polar solvents. The solute–solute and solvent–solvent forces are dispersion forces and are of comparable strengths to the solute–solvent forces found in the solution.

Ionic substances are not soluble in non-polar or slightly polar solvents because the energy to be gained in solute–solvent forces is much less than that required to break the solute–solute (ionic) bonds.

In the experiment you will investigate the solubility of an ionic compound (CuSO_4) and a molecular substance (I_2) in three solvents: water (hydrogen bonded), acetone (polar), and cyclohexane (non-polar, structure shown on right). You will also observe the miscibility of a polar, hydrogen bonded liquid (propan-2-ol) with water and a non-polar solvent (cyclohexane).



The miscibility of two liquids can be affected by a solute dissolved in one of them. For example, if a molecular liquid is miscible with water, adding a soluble ionic substance to the mixture can reduce the solubility of the molecular liquid by forming strong bonds to the water molecules. This process is called **salting out** and is commonly used to reduce the solubility of a liquid in water.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
solid $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$		
solid $(\text{NH}_4)_2\text{SO}_4$		
solid I_2		
propan-2-ol		
acetone		
cyclohexane		
food dye		

APPARATUS	RISK	SAFETY PRECAUTIONS
All glassware including beakers, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Do not use damaged glassware. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Materials Needed

semi-micro test tubes and rack	solid I ₂	propan-2-ol	food dye
1.5 × 15 cm test tubes and rack	acetone	solid CuSO ₄ ·5H ₂ O	solid (NH ₄) ₂ SO ₄
squash pipettes	cyclohexane		

Method

Part A: Computer Simulation of Vapor Pressure



This part of the experiment will be performed as a demonstration in the lab by the lab supervisor or your demonstrator.

The simulator allows the user to choose a substance, then shows the vapor pressure being recorded as the substance vaporizes in a flask connected to a mercury manometer. The vapor pressure of any substance varies with temperature and the simulator chooses a temperature at which the vapor pressures will be measured when it is started. Because the simulator accurately portrays the measurement process, it may take up to 30 seconds for the liquid–vapor equilibrium to be established.

Part B: Solubility of ionic and molecular solids

- Place a small amount (about the size of 1 grain of rice, see picture) of copper sulfate into each of three DRY semi-micro test tubes. Add 20 drops of water to the first test tube and gently flick the test tube with your finger to ensure mixing.
- Repeat step 1 using acetone in place of water as the solvent in the second test tube.
- Repeat step 2 replacing acetone with cyclohexane in the third test tube.
- Hold the test tubes against a white background to compare the solubility of copper sulfate in the three solvents and record your results. Discard organic solvents into the organic waste container in the fume cupboard. Once these test tubes have been emptied you can re-use them in the next step without further cleaning.
- Repeat steps 1–4 but replace the copper sulfate with a tiny piece of solid iodine (I₂) in each of the test tubes.



Part C: Miscibility of liquids

- Place 10 drops of propan-2-ol into each of two DRY semi-micro test tubes.
- Add 10 drops of water to one test tube and 10 drops of cyclohexane to the other test tube, shake and observe whether the two liquids in each test tube are miscible. Record your results then discard the contents of the test tubes into the organic waste container.
- Place 10 drops of water in a clean semi-micro test tube and add 10 drops of cyclohexane to it. Record whether the two liquids are miscible. If two layers form (*i.e.* the liquids are immiscible) add a drop of blue food dye to the mixture and shake, then allow the two layers to reform. Based on your observation, decide whether the food dye is more soluble in water or cyclohexane and then discard the mixture into the organic waste container.

Part D: Salting out of one liquid from a mixture

- Place 30 drops of propan-2-ol in a 1.5 × 15 cm test tube. Add 1 drop of food dye solution and 2.5 mL of water and shake. What you see should be consistent with your observation of the miscibility of water and propan-2-ol from part C above. Record the appearance of this mixture in the results section.
- Add a small amount of solid ammonium sulfate, about enough to cover a 5 cent coin. Shake and let the solid dissolve. If the solid dissolves rapidly, add a second similar amount and shake to allow it to dissolve. If the second portion still dissolves rapidly, add a third similar amount.
- Let the test tube stand for one minute, then record the appearance of the contents of the test tube.
- Dispose of the mixture in the organic residues container.
- Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.

School of Chemistry, UNSW

Coversheet for Submission of Individual Report

Report: Intermolecular Forces

Circle your course: CHEM1011 CHEM1031 CHEM1051

Student Name (full name): _____

Student Number (e.g., z1234567): _____ Lab. class: day/time: _____

Demonstrator: _____ Workspace: _____

Date and Time Submitted: _____

In preparing this assessment task I have followed the Student Code of Conduct. I certify that I have read and understand the University requirements in respect of student academic misconduct outlined in the Student Code of Conduct and Annexure 1 of the Student Misconduct Procedures. I declare that this assessment item is my own work, except where acknowledged, and has not been submitted for academic credit previously in whole or in part. I acknowledge that the assessor of this item may, for assessment purposes: (1) provide a copy to another staff member of the University; (2) communicate a copy of this assessment item to a plagiarism checking service (such as Turnitin) which may then retain a copy of the assessment item on its database for the purpose of future plagiarism checking. I have retained a copy of this, my assignment, which I can provide if necessary. By signing this declaration I am agreeing to the statements and conditions above.

Student signature and date: _____

Full UNSW policy concerning student misconduct is available at: <https://student.unsw.edu.au/conduct>

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Non-core Skills

Applying chemical principles: correct logic and explanation used appropriate principle(s)

Describing chemical changes: correct use of terms, e.g., precipitate, supernatant

Mastery (Applying chemical principles): clear & concise explanation(s)

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

RESULTS

Part A: Computer Simulation of Vapor Pressure

- Record the temperature and then the vapor pressure for each substance. Temperature = °C
- Indicate the type(s) of intermolecular force acting between the molecules for each substance.

methanol CH ₃ OH Vapor pressure:	ethanol CH ₃ CH ₂ OH Vapor pressure:	propan-1-ol CH ₃ CH ₂ CH ₂ OH Vapor pressure:
<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding

Briefly discuss the trend in vapor pressure for these substances in terms of intermolecular forces:

Substance with highest normal boiling point: _____

2-chloropropane H ₃ C — C(Cl) — CH ₃ Vapor pressure:	propan-2-ol H ₃ C — C(OH) — CH ₃ Vapor pressure:	propanone H ₃ C — C(=O) — CH ₃ Vapor pressure:
<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding

Briefly discuss the trend in vapor pressure for these substances in terms of intermolecular forces:

Substance with highest normal boiling point: _____

pentane CH ₃ CH ₂ CH ₂ CH ₂ CH ₃ Vapor pressure:	hexane CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃ Vapor pressure:	heptane CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃ Vapor pressure:
<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding	<input type="checkbox"/> Dispersion <input type="checkbox"/> Dipole/dipole <input type="checkbox"/> Hydrogen bonding

Briefly discuss the trend in vapor pressure for these substances in terms of intermolecular forces:

Substance with highest normal boiling point: _____

Part B: Solubility of ionic and molecular solids

Complete the following table with regard to the solubility of each substance in the various solvents.

	Water	Acetone	Cyclohexane
Copper sulfate			
Iodine			

Questions

1. What type or types of solvent are ionic substances usually soluble in?

2. What type or types of solvent are molecular substances usually soluble in?

Part C: Miscibility of liquids

Record whether the combinations of liquids shown below are miscible.

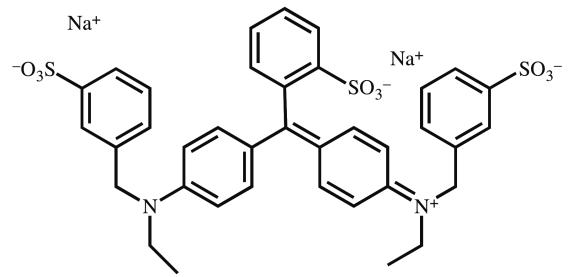
	Water	Cyclohexane
Propan-2-ol		
Cyclohexane		

Questions

1. What did you observe when food dye was added to the mixture of water and cyclohexane?

2. What can you deduce from this observation about the forces between the food dye and water and cyclohexane?

The chemical structure of the blue food dye (Brilliant Blue 133) is shown on the right. This is an abbreviated structure and carbon and hydrogen atoms are not explicitly shown, however the bonds between carbon atoms and all other atoms apart from hydrogen are shown.



3. What aspect of the chemical structure of the dye is responsible for its relative solubility in water compared to cyclohexane?

Part D: Salting out of one liquid from a mixture

Describe the appearance of the mixture of propan-2-ol, water and food dye at step 1.

Describe the appearance of the mixture of propan-2-ol, water and food dye after the addition and dissolution of the ammonium sulfate.

Question (part D)

Explain your observations in terms of the new intermolecular forces introduced when the ammonium sulfate was dissolved in the mixture.

Intermolecular forces and shapes of molecules [For CHEM1031 and CHEM1051 only]

From the first part of this experiment, write down the vapor pressures of propan-1-ol and propan-2-ol.

propan-1-ol: _____ propan-2-ol: _____

What term describes the relationship between these two molecules? Be as specific as you can.

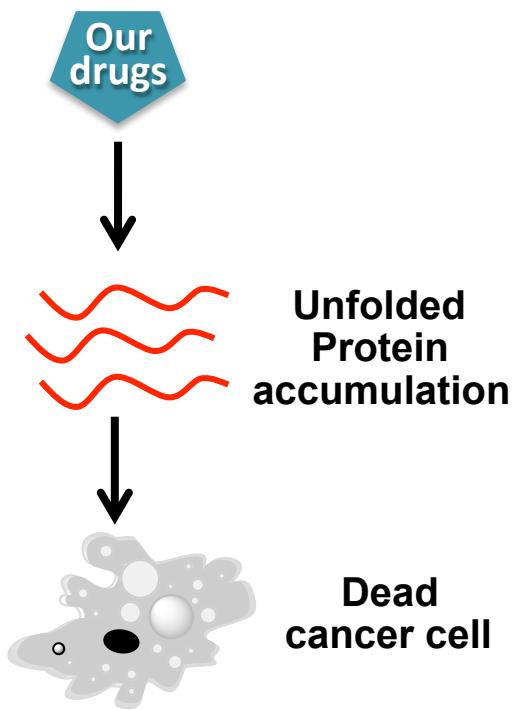
Propan-1-ol and propan-2-ol are: _____

In no more than two sentences explain the difference between the vapor pressures of these two substances in terms of intermolecular forces.

Intermolecular forces and the structures of proteins [For CHEM1051 only]

Proteins fold into complicated shapes determined by many intermolecular forces within the molecule (hydrogen bonds, interactions of charged groups, dipole-dipole forces). If the natural shape of a protein is changed (the protein is then said to be denatured) the protein usually will not perform its biological functions. Concentrated solutions of inorganic salts can denature proteins. How might this happen? (Hint: water molecules can be important in maintaining the natural shape of a protein.)

Cool Research in the School of Chemistry



A/Professor Shelli R McAlpine and her group are making new drugs that target cancer.

The research in my group is aimed at discovering new molecules that kill cancer cells. We make organic molecules (e.g. those made from carbon, oxygen, nitrogen, and sulphur) and test them for their ability to kill cancer cells. We have discovered several molecules that stop proteins from being folded correctly. The build up of unfolded proteins in cancer cells produces cell death.

Projects are primarily synthetic, where you would make molecules that target proteins involved in the protein folding process within the cells. Experiments utilise organic synthetic techniques including setting up and running reactions in a solution and on solid-phase. The projects also involve analysis of products, learning purifications techniques, and producing a final compound that will be tested in our biological assays.

Synthesis of molecules that target heat shock protein 90 (Hsp90)

Cancer is still considered an incurable disease. We have recently developed molecules that block a protein that enables cancer cells to grow. This protein, heat shock protein 90 (Hsp90), controls the protein folding pathways of most new proteins produced by cancer cells. Using our molecule we stop Hsp90 from folding proteins, which leads to unfolded proteins accumulating in the cell. Protein accumulation leads to cell death because there are not enough working proteins required for the cell to function. Making molecules on this project will provide data on which structures are most effective at blocking cell growth.

Synthesis of molecules that target heat shock protein 70 (Hsp70)

Heat shock protein 70 (Hsp70) helps Hsp90 to fold proteins. Hsp70 also protects the cells from stress mechanisms and helps cancer cells survive. We have discovered a molecule that stops Hsp70 from helping the cell survive stress and blocks its protein folding functions. By blocking Hsp70 and its protection mechanisms, cancer cells die. This project involves making molecules that bind selectively to Hsp70.

Synthesis of molecules that target heat shock protein 27 (Hsp27)

Heat shock protein 27 (Hsp27) is the first protein involved in Hsp90's protein-folding cascade. There are no molecules that block Hsp27's function. Our goal is to develop organic molecules that bind to and inhibit the function of Hsp27 and to test their ability to stop protein folding and cancer cell growth.

If you are interested in a 3rd year, Honours or Ph. D. project in this area drop me an email (s.mcalpine@unsw.edu.au), tour the lab, and talk to our students and post-docs.

CHEMICAL EQUILIBRIUM

Aim

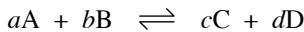
The aim of this experiment is to study the effect of external changes (e.g., addition of reactants or products, changes in temperatures) upon chemical systems at equilibrium.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

In this series of experiments, equilibrium will be considered in a qualitative fashion. Parts B, C, and D of the experiment illustrate the dynamic nature of equilibrium, and the predictable manner in which equilibrium systems react to changes as described by Le Chatelier's principle.

Chemical equilibrium occurs when two opposing reactions take place at the same rate. Although no chemical change is visible, the system is in dynamic equilibrium. For the generic reaction:



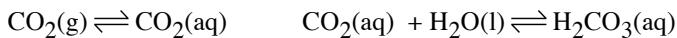
the ratio of product concentrations to reactant concentrations (all expressed as mol L⁻¹) at a given temperature is a constant, represented by K_c , the equilibrium constant with respect to concentration.

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Le Chatelier devised a principle which applies to *systems at equilibrium*: if a system at equilibrium is disturbed (e.g., by heating, by adding more of a reactant or a product, or by removing some of a reactant or product) the system will change (adjust itself) so as to minimise the disturbance and return to equilibrium again. Experimental observations on many chemical systems have led to the generalisation that all systems undergoing a chemical reaction move towards equilibrium.

In this assignment some systems at equilibrium are examined and the conditions under which the equilibrium position can be changed are also examined. Any changes in the equilibrium position will be detected by noting changes in gas volume, colour changes, and/or the formation or disappearance of a precipitate.

In part A you will look at the equilibrium between gaseous and dissolved carbon dioxide and carbonic acid formed by the reaction of the dissolved CO₂ with water:



Carbonic acid is a diprotic weak acid and dissociates in two steps:

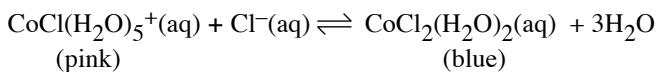


These equilibria can be disturbed by adding $\text{OH}^-(\text{aq})$ which reacts with the $\text{H}^+(\text{aq})$ to form water resulting in a cascade of changes which ultimately changes the amount of carbon dioxide which dissolves.

In part B, equilibria involving CO_2 , HCO_3^- and CO_3^{2-} will be studied.

Part C investigates the equilibrium which exists between $\text{Cr}_2\text{O}_7^{2-}$ and CrO_4^{2-} and its dependence upon the pH of the solution and in part D this is extended by the incorporation of a second equilibrium involving Ba^{2+} and CrO_4^{2-} .

The value of the equilibrium constant for a reaction usually varies with temperature. For endothermic reactions (taking in heat; $\Delta H > 0$) as the temperature increases, so does the equilibrium constant. For exothermic reactions (producing heat; $\Delta H < 0$) the equilibrium constant decreases as the temperature increases. In part E you will investigate this for the reaction:



Pre-lab Questions (compulsory, must be completed to be allowed to start the experiment)

What colour is the chromate anion (CrO_4^{2-})?

What colour is the dichromate anion ($\text{Cr}_2\text{O}_7^{2-}$)?

Is barium chromate soluble or insoluble in water?

Is barium dichromate soluble or insoluble in water?

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

The web page where you get your safety information has a link to a VIDEO about one of the techniques you will use in this experiment. **You should watch this video before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
2 M NaOH		
saturated $\text{Ca}(\text{OH})_2$		
dry ice (solid CO_2)		
1 M NaHCO_3		
0.1 M $\text{K}_2\text{Cr}_2\text{O}_7$		
0.1 M K_2CrO_4		
2 M HNO_3		
universal indicator		
0.1 M $\text{Ba}(\text{NO}_3)_2$		

CoCl ₂ ·6H ₂ O		
propan-2-ol		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Materials Needed

2 M NaOH	0.1 M K ₂ CrO ₄	propan-2-ol	100 or 150 mL beakers
saturated Ca(OH) ₂	2 M HNO ₃	side-arm test tube with plastic tube	steam bath
1 M NaHCO ₃	0.1 M Ba(NO ₃) ₂	10 mL measuring cylinder	ice
0.1 M K ₂ Cr ₂ O ₇	Universal indicator	test tubes	cotton wool
dry ice	CoCl ₂ ·6H ₂ O	test tube rack	small plastic funnel



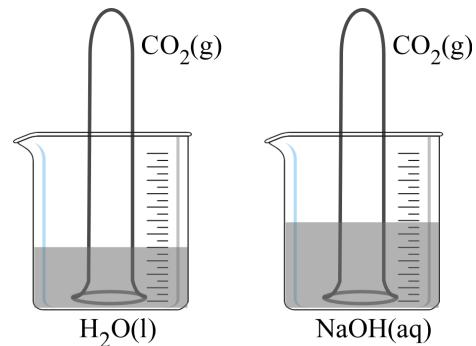
DRY ICE WILL FREEZE SKIN

Avoid skin contact with dry ice – it can cause severe skin damage. Use metal tongs to pick up pieces of dry ice. DO NOT LEAVE THE TONGS IN THE DRY ICE CONTAINER.

Part A - Reactions of Carbon Dioxide Gas

Method

- Place a small lump of dry ice into a large test-tube. To the dry ice add about 5 mL of water. After the gas has bubbled through the water for a few minutes (so that it may be assumed the solution is saturated with the gas), test the acidity of the solution by adding 1 drop of Universal indicator solution and check the pH on the colour chart provided. Record your observations onto your results page.
- Place two test tubes in the test tube rack. In one tube place a small lump of dry ice and cover the mouth of the tube with a loose plug of cotton wool. Stand the test tube in a beaker of containing a little warm water (from the hot water tap) for a minute. The test tube will fill with carbon dioxide gas. In the meantime, obtain two small (100 or 150 mL) beakers, and place 20 mL of water in each and add 3 drops of universal indicator to each.
- Do this step quickly to avoid loss of CO₂(g) from the test tube. Remove the cotton wool and tilt the test tube to allow the dry ice to slide into the second test tube and invert the first tube into one of the small beakers containing water (see adjacent figure). Transfer the cotton wool plug to the second test tube which should now have the solid CO₂ in it.
- Warm the second tube in warm water for a minute, then remove the cotton wool, let the CO₂(s) slide out into an empty beaker and invert the tube into the second small beaker. Add 20 mL of 2 M NaOH to the water in the second beaker and move the test tube around to stir the solution



(see adjacent figure).

- Observe both test tubes over a 10 minute period, noting any changes in the liquid level in the two test tubes. Record your observations.

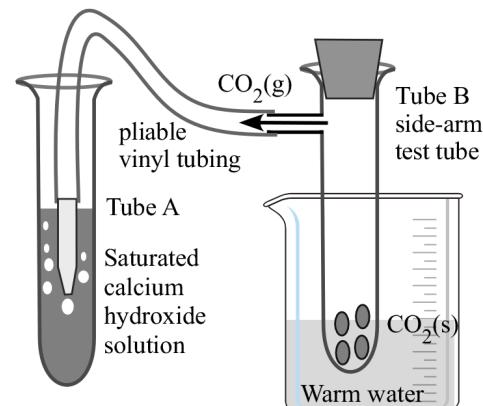
Part B – Carbonate Equilibrium

Method



CAUTION 2 M HNO₃ is corrosive – be careful not to splash or spray the solution during rinsing in the next step.

- Choose a wide test tube from your locker to use as tube A in the apparatus shown on the right. Add 2 M HNO₃ to the test tube to a depth of about 2 cm, then add tap water to half-fill the tube. Let it soak with a little stirring for 30 seconds, then empty the acid into the sink. Rinse the test tube with LOTS OF WATER (rinse under running water in the sink).
- Pour about 20 mL of saturated Ca(OH)₂ solution into this test-tube (tube A in diagram on right). If the saturated Ca(OH)₂ solution is cloudy you may need to filter it into the tube. Ask your demonstrator if filtration is required.
- Place several small pieces of dry ice into test-tube B. Fit the cork into test tube B and insert the end of the plastic tube into test tube A.
- Place tube B into a beaker about half full of warm water from a hot water tap (do not use a water bath for this, it is too hot). When you see gas bubbling through tube A, as shown in the diagram, record the **immediate change** you see onto your results page. Ask your demonstrator for assistance if gas does not bubble through the liquid in tube A.
- Continue passing CO₂ into the solution and observe any further changes for about 5 minutes. Record any **second observable change** in your result page next to the heading 'Addition of excess CO₂(g)'. Consult your demonstrator if you do not see a second change to the solution in tube A. *Remove the plastic tube from tube A.*
- If the solution in tube A is cloudy or turbid, filter it through a small plug of cotton wool in a plastic funnel into a clean test tube and use the filtered solution in the following step.
- Pour one third the contents of tube A into a clean test-tube (labeled tube C) and pour one third the contents of tube A into a clean test-tube (labeled tube D). The remaining solution in tube A serves as a 'control' to compare with the other tubes.



The steam bath, hot water, and steam can all cause severe burns.

Do not touch the steam bath. Do not reach into the steam or touch the water.

- Heat the solution in test-tube C on the steam bath (hold the tube in a wooden test-tube holder) for a few minutes. Compare the solution in test-tube C to that in tube A. Record your observations on the results page. Explain your observations in terms of the equilibria that exist between the carbonate ion, the hydrogencarbonate ion and calcium carbonate.
- Add 1 M NaHCO₃ dropwise (a maximum of 20 drops) to test-tube D and compare the contents of test-tube D to that of tube A. Record your observations on the results page. Explain your observations in terms of the equilibria that exist between the carbonate ion, the hydrogencarbonate ion and calcium carbonate.
- Carefully rinse all test tubes with a few mL of 2 M HNO₃ followed by plenty of water. **Remember that 2 M nitric acid is corrosive so do not splash or spill it.**



Dispose of all waste solutions containing chromium into the appropriate waste container in the fumecupboard.

Part C – The Chromate/Dichromate Equilibrium

Method

1. Number four test-tubes from 1 to 4.
2. Pour 0.1 M K_2CrO_4 (2 mL) into tubes 1 and 2, and 0.1 M $\text{K}_2\text{Cr}_2\text{O}_7$ (2 mL) into tubes 3 and 4. Observe the colours of these solutions and record them in the results sheet.
3. Add 2 M HNO_3 (2 mL) to tube 1 and tube 3. Record any colour changes observed.
4. Add 2 M NaOH (2 mL) to tubes 2 and 4. Record any colour changes observed.
5. Answer the questions contained in the results page relating to this equilibrium system.

Part D – The Chromate/Dichromate Equilibrium with Ba^{2+}

Method

1. Place 0.1 M K_2CrO_4 (20 drops) in a clean test-tube. Add 2 M NaOH (5 drops). Add, a drop at a time, 0.1 M $\text{Ba}(\text{NO}_3)_2$ **until you notice a change**. If you get to 10 drops added without any observable change, consult your demonstrator. Record your observations on the results page.
2. To this test-tube add drop by drop 2 M HNO_3 until a change is noted. Record your observations.
3. Place 0.1 M $\text{K}_2\text{Cr}_2\text{O}_7$ (20 drops) in a clean test-tube. Add 2 M HNO_3 (5 drops). Add, a drop at a time, 0.1 M $\text{Ba}(\text{NO}_3)_2$ to a maximum of 10 drops. Record in pen your observations on the results page.
4. To this test-tube add, drop by drop, 2 M NaOH until a change is noted. Record your observations.
5. Rinse all test tubes with 2 mL of 2 M HNO_3 followed by plenty of water.
6. Complete all equations.

Part E – Variation of the Equilibrium Constant with Temperature

Method

1. Clean a 1.5×15 cm (tall, narrow) test tube with a bottle brush, detergent and water. Rinse the tube with water then drain as much water as possible from it. Add about 2 mL of propan-2-ol, swirl it around the inside of the tube and discard it into the organic waste container. This should remove most of the water from the tube.
2. Weigh 0.1 g of $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ onto a watch glass.
3. Transfer the solid to the test tube and add 2.5 mL of propan-2-ol. Swirl the tube until most of the cobalt chloride dissolves to produce a blue solution.
4. Add water one drop at a time, swirling the tube between drops, until the solution turns purple (not pink). If the solution turns pink, add drops of propan-2-ol with swirling until it turns purple.
5. Cool the test tube in an ice bath. Record the colour of the cold solution in the appropriate location in the results table.
6. Heat the tube in a steam bath in the fume cupboard. Record the colour of the hot solution.
7. Discard the solution into the metals waste container.
8. Based on your observations, deduce whether the reaction is exothermic or endothermic.
9. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room. Ask your demonstrator to check your locker.

Blank page.
(Can be used for working or calculations)

School of Chemistry, UNSW

Coversheet for Submission of Individual Report

Report: Chemical Equilibrium

Circle your course: CHEM1011 CHEM1031 CHEM1051

Student Name (full name): _____

Student Number (e.g., z1234567): _____ Lab. class: day/time: _____

Demonstrator: _____ Workspace: _____

Date and Time Submitted: _____

In preparing this assessment task I have followed the Student Code of Conduct. I certify that I have read and understand the University requirements in respect of student academic misconduct outlined in the Student Code of Conduct and Annexure 1 of the Student Misconduct Procedures. I declare that this assessment item is my own work, except where acknowledged, and has not been submitted for academic credit previously in whole or in part. I acknowledge that the assessor of this item may, for assessment purposes: (1) provide a copy to another staff member of the University; (2) communicate a copy of this assessment item to a plagiarism checking service (such as Turnitin) which may then retain a copy of the assessment item on its database for the purpose of future plagiarism checking. I have retained a copy of this, my assignment, which I can provide if necessary. By signing this declaration I am agreeing to the statements and conditions above.

Student signature and date: _____

Full UNSW policy concerning student misconduct is available at: <<https://student.unsw.edu.au/conduct>>

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Core Skills

Heating materials: steam bath – correct use (last chance)

Non-core Skills

Applying chemical principles: correct logic and explanation used appropriate principle(s)

Chemical equations: balanced - atoms & charge correct formulae correct states of matter no spectator ions

Describing chemical changes: correct use of terms e.g., precipitate, supernatant

Mastery (Applying chemical principles): clear & concise explanation(s)

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: in pen (no whiteout) in lab manual

Safety awareness: safety eyewear on at all times

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

RESULTS

(Remember to not include spectator ions in ionic equations.)

Part A – Reactions of Carbon Dioxide Gas

Observations at end of step 1.

Is an aqueous solution of CO₂ acidic, basic or neutral? Include an equation to illustrate this behaviour.
(e.g., if acidic, the equation will have H⁺(aq) as a product and if basic, the equation will have OH⁻(aq) as a product.)

Solubility of Carbon Dioxide

Observations at end of steps 2–5.

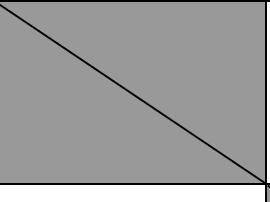
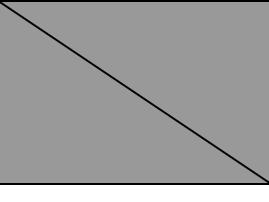
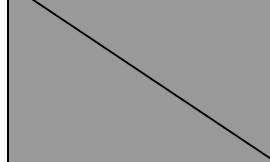
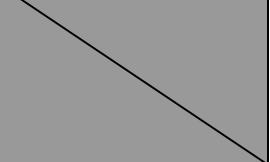
Compare the solubility of CO₂ in H₂O and in aqueous NaOH. Explain any differences which you observe, including any relevant equations.

Part B - Carbonate Equilibrium

PROCEDURE	OBSERVATIONS	EQUATIONS
Initial bubbling of CO ₂ (g) (STEP 4) (immediate change)		
Addition of excess CO ₂ (g) (STEP 5) (second observable change)		
Heating tube C (STEP 8)		
Adding 1 M NaHCO ₃ to tube D (STEP 9)		

Explain your observations in terms of the equilibria involved.

Part C - The Chromate/Dichromate Equilibrium

	Tube 1	Tube 2	Tube 3	Tube 4
Initial Colour				
Colour after adding HNO_3				
Colour after adding NaOH				

Questions

1. What do you conclude about the dependence of the conversion of chromate to dichromate upon $[\text{H}^+]$?
Write a chemical equation to illustrate this behaviour.

2. What do you conclude about the dependence of the conversion of dichromate to chromate upon $[\text{OH}^-]$?
Write a chemical equation to illustrate this behaviour.

Part D - The Chromate/Dichromate Equilibrium with Ba²⁺

	Initial observation after adding Ba ²⁺	Observation on dropwise addition of:	
		HNO ₃	NaOH
Tube from steps 1 and 2			
Tube from steps 3 and 4			

Write balanced ionic equations for the chemical reactions which you have observed.

Tube from Steps 1 and 2.

Initial reaction:

Ionic equation	
----------------	--

Reaction after adding HNO₃:

Ionic equation	
----------------	--

Tube from Steps 3 and 4.

Initial reaction:

Ionic equation	
----------------	--

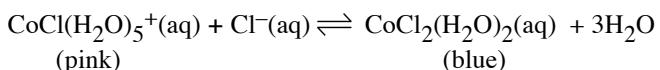
Reaction after adding NaOH:

Ionic equation	
----------------	--

Part E – Variation of the Equilibrium Constant with Temperature

Colour of solution above room temperature	
Colour of solution at room temperature	purple
Colour of solution below room temperature	

Is the reaction below exothermic or endothermic, and how can you deduce this?

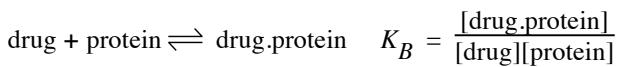


The above reaction is: _____

because:

Drug–Target Binding [For CHEM1051]

The binding of a drug to a target protein can be quantified by an equilibrium constant, the association or binding constant.



It is the concentration of the drug-protein complex which determines the medicinal effects of the drug.

1. Why does increasing the dose of a drug usually increase its medicinal effect?

2. Effective drugs usually have K_B values of 10^6 or larger. Why are such large values for K_B desirable?

ACID – BASE TITRATION

Aim

The aims of this experiment are to become familiar with the use of pH meters and to develop an understanding of pH and weak acid/base equilibria by constructing the titration curve for a weak acid or weak base.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

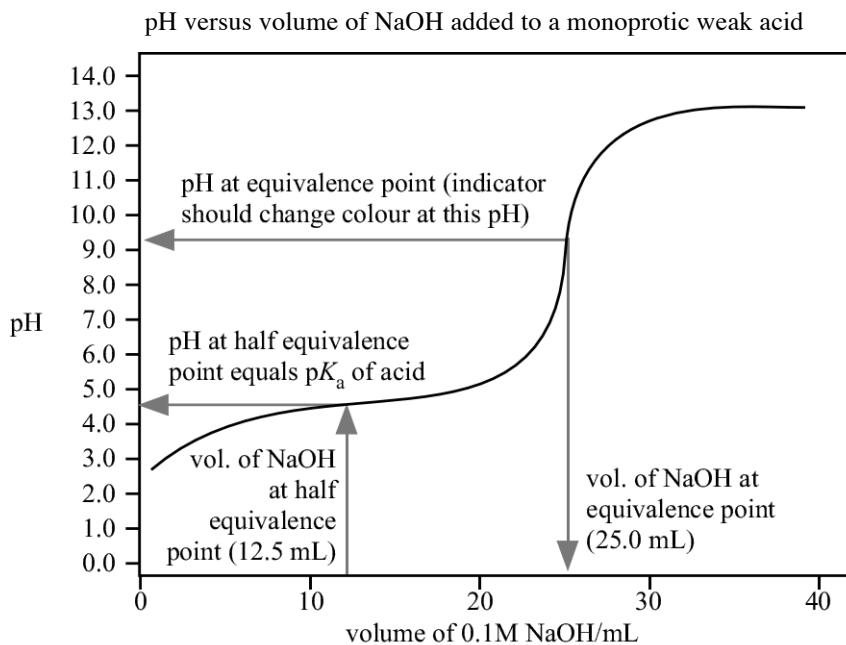
Introduction

The acidity or alkalinity of a solution can be expressed in terms of the pH of the solution, defined by:

$$\text{pH} = -\log_{10}([\text{H}^+])$$

where the H^+ concentration is in mol L⁻¹ (more correctly $\text{pH} = -\log_{10}([\text{H}^+]/m^0)$ where $m^0 = 1 \text{ mol/L}$). The lower the pH, the more acidic the solution; the higher the pH, the more alkaline the solution. At 25 °C a neutral solution has a pH of 7.

When an acid is titrated with a base, the pH is initially low. As more base is added, the acid is consumed and the pH begins to rise, slowly at first. As the equivalence point of the titration is approached, most of the acid has been consumed and the pH begins to rise more sharply. The equivalence point is marked by the steepest rise in pH as base is added (mathematically the equivalence point is an inflection in the titration curve). Beyond the equivalence point, the base is in excess and as more base is added, the pH change slows and eventually changes very little as more base is added. If the pH of the solution is measured throughout a titration, a titration curve showing pH as a function of the volume of added base can be sketched as shown below.



For a weak monoprotic acid (written as HA below), an equilibrium exists in aqueous solution between it and its conjugate base, A^- (for example, HA might be acetic acid and then A^- would be the acetate anion) and there is an equilibrium constant (the acid dissociation constant or K_a) for this equilibrium.



At the start of the titration, before any base has been added, the weak acid is present mostly in the form HA. As the titration proceeds, the concentration of HA decreases and the concentration of A^- increases. At the equivalence point of the titration, sufficient OH^- has been added to react with all the HA and the weak acid is now present almost entirely as its conjugate base, A^- . At a point half way between the start and the equivalence point, enough base has been added to react with half the acid so half the acid has been converted into its conjugate base, and $[\text{HA}] = [\text{A}^-]$.

Substituting into $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$ gives $K_a = [\text{H}_3\text{O}^+]$ which means that $\text{p}K_a = \text{pH}$ at the 'half equivalence point'.

Thus at the half equivalence point in a titration, the pH equals the $\text{p}K_a$ which can then be read from the titration curve.

The solution at the half equivalence point is a buffer (since $[\text{HA}] = [\text{A}^-]$). In a buffer, small amounts of H^+ or OH^- can be added without significant change in the pH of the solution, and this can be seen from the graph above.

If a weak base is titrated with a strong acid, the titration curve will start at a high pH, moving to a lower pH as the titration proceeds. The principles discussed above about titration curves still hold. Even though it is a base which is being titrated, the pH at the half equivalence point still represents a K_a – this time of the conjugate acid of the base being titrated.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

The web page where you get your safety information has a link to VIDEOS about several of the techniques you will use in this experiment. **You should watch these videos before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
0.1 M HCl		
0.1 M NaOH		
methyl red		
phenolphthalein		
0.1 M unknown acid or base		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Materials Needed

50.0 mL burette	conical flask	pH meter with pH 4 buffer	methyl red
25.0 mL pipette pipette filler	small (100 or 150 mL) beaker red and blue litmus paper	standardized 0.1 M HCl standardized 0.1 M NaOH	phenolphthalein 0.1 M unknown acid or base

Techniques

Read the instructions below and watch the video online (link on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these techniques.

Technique – Using a pH meter

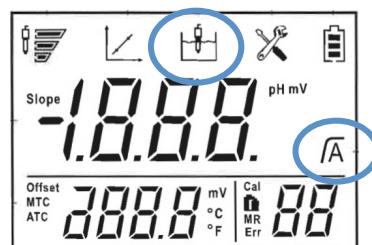
There are two types of pH meter in use in the laboratory. Follow the instructions for the type of meter you have on your bench. The pH meter has been calibrated before the laboratory class. **You do not need to calibrate it. If you believe the pH meter is not working correctly, ask the laboratory staff to check it.**

When measuring the pH of the solutions, keep the solution stirred by gently moving the beaker around. The pH electrode is fragile, do not use it to stir the solutions.



Mettler Toledo meters: Place the electrode in the solution. If the 'A' symbol (right side of display) is shown, press and hold the green Read button until the 'A' disappears. The meter will now continuously display the pH of the solution.

The Measurement icon (top middle) should be visible all the time you are measuring pH values.



Hanna ('stick') meters: Place the electrode in the solution and read the pH from the meter.

Technique – Drawing scientific graphs

A scientific graph should have a **title** in the form of '*y axis quantity* versus *x axis quantity* for *substance/reaction*'. For example, if you measured the temperature of a mixture of water and sodium chloride over time you would label the graph of your data as 'Temperature versus time for sodium chloride dissolving in water.' If you measured the solubility of calcium hydroxide at different temperatures your graph would be titled 'Solubility versus temperature for calcium hydroxide.' In some cases the *x*-axis name can include the name of the substance for which the data was collected, *e.g.* 'Absorbance versus $[MnO_4^-]$.'

Axis labels should be in the form of 'quantity/units' for example 'T/K', 'time/s', ' $[Cu^{2+}]$ /mol L⁻¹'.

Scales should be chosen so that the data fill the area of the graph. If the data are clustered away from the origin (0,0), then the origin should not be included on the graph, unless it is desired to show that the line of best fit through the data passes through the origin.

Data points must be plotted clearly. Data points (and ONLY data points) may be plotted using pencil.



At the start of the experiment the pH electrode will be standing in a buffer solution of pH 4 in a conical flask. DO NOT CONTAMINATE THIS SOLUTION.

Do not throw out the solution. At the end of the experiment wash the electrode using a wash bottle, allow it to drain and return it to this solution.



Do not waste the standardised acid and base solutions. These solutions take considerable time and effort to prepare. You need about 10 mL of solution to rinse your burette and no more than 100 mL total for the titrations.

Method

1. The unknown is provided in dispensers. Obtain no more than 60 mL of the unknown assigned to your workspace into a **clean, dry, labelled** beaker. Determine whether it is an acid or a base by placing a drop on each of red and blue litmus papers using a clean stirring rod. **Do not put litmus paper into the unknown solution.**
2. (Titration #1, with indicator only) Pipette 25.0 mL of the unknown into a conical flask. Add 2–3 drops of either phenolphthalein (if the unknown is an acid) or methyl red (if the unknown is a base) and titrate with either standardised base (~0.1 M NaOH) or standardised acid (~0.1 M HCl) in order to determine the volume required to reach the end point. There is no need to repeat this titration. The **horizontal axis on your graph** for the titration using the pH meter (below) should go from zero to the endpoint determined in this step plus 10 mL (rounded up to the next whole 5 mL).



Plot as you go and keep the electrode in the solution during the titration.

Plot the titration curve on the graph as each measurement is taken. Reduce the volume of solution added when the pH starts to change more rapidly (*i.e.* the graph starts to curve).

3. (Titration #2, with pH meter, but no indicator) Pipette 25.0 mL of the unknown solution into a **small (100 or 150 mL) beaker**. Remove the electrode from the buffer solution, and rinse the electrode with water from your wash bottle. Lower the electrode into the solution and measure and record its pH. **Make sure the bulb on the end of the electrode is completely covered by the solution.** Starting with the burette filled to the 0.0 mL mark, add 1.0 mL of the appropriate standard solution (from the burette) and measure the pH of the solution. Record the volume and pH on the results sheet and plot the value on your graph.
4. Keep the electrode in the solution (with the bulb covered by the solution) until you have finished the titration. Do not remove or rinse the electrode. Continue adding 1.0 mL aliquots of standard solution from the burette to the unknown, recording the pH value and plotting the value after each addition. As soon as the pH starts to change more rapidly (that is, your plot becomes more curved), decrease the additions of standard solution to 0.50 mL. **Record and plot** the pH after each addition.
5. Continue the titration with 0.50 mL additions until a volume of about 3 mL after the equivalence point (sudden change in pH) is reached. Then proceed with 1.0 mL additions until a volume of at least 10 mL beyond the endpoint (where the indicator changed colour in the first titration) is reached. Record the pH value and plot each value as you go.
6. At the completion of the titration, remove the electrode from the solution, rinse thoroughly with water from your wash bottle, and return it to the pH 4 buffer solution in the conical flask where it was when you started the experiment.
7. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.
8. From the plot of pH versus volume of titrant added, determine pK_a of the acid unknown or of the conjugate acid of the base unknown, as appropriate.

Blank page.
(Can be used for working or calculations)

RESULTS (Results must be written in pen, only graphs may be drawn in pencil)

Your name: _____ Lab. Day/ Time: _____ Workspace: _____

Data for Titration Curve

Unknown number

ANSWER

Acid or base? Acid

Acid

Base

Concentration of standard acid or base: mol/L

Type of pH meter (Toledo or Hanna):

Preliminary (first) Titration

Burette reading at start: mL

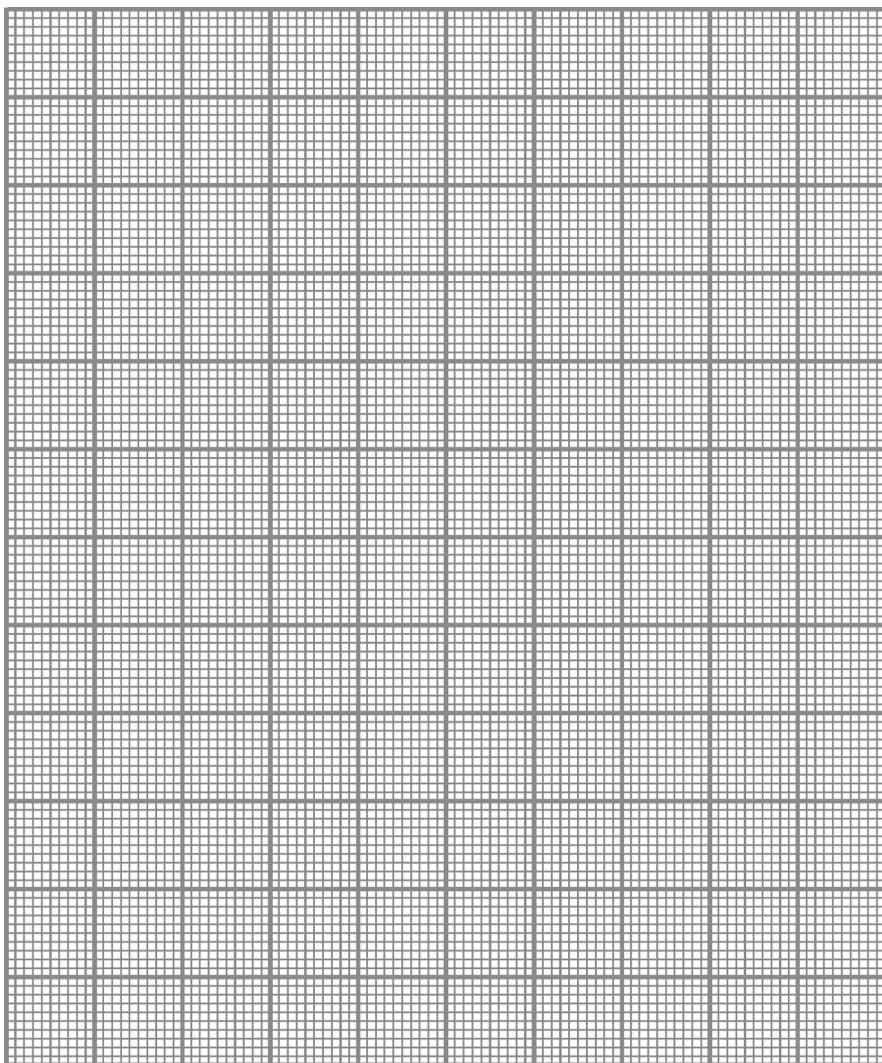
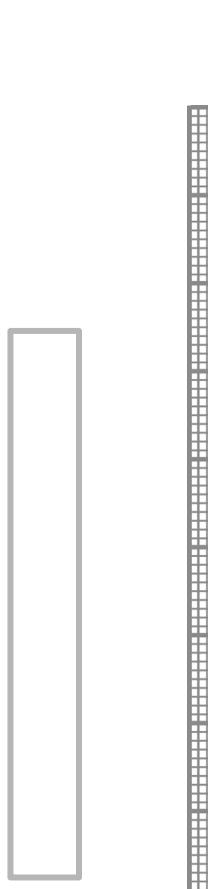
Burette reading at endpoint: mL

Titre: mL

Titration Curve (second titration)

As you record the results, plot the titration curve.

Plot your titration curve below using the data you have recorded. Give the graph an appropriate title and label the axes appropriately. Indicate on the graph the equivalence point and the half equivalence point.



p*K_a* of (circle the correct term): weak acid OR conjugate acid of the weak base

Suggest a suitable indicator which could be used to carry out the titration (refer to SI Chemical Data).

Recommended indicator:

Explain why you chose this indicator.

Acids and Bases in Biochemistry [For CHEM1031 and CHEM1051 only]

Amino acids contain a basic amine group ($-\text{NH}_2$) and a carboxylic acid group ($-\text{COOH}$), e.g. glycine, $\text{NH}_2\text{CH}_2\text{COOH}$. In acidic solutions both the amine and the acid group are protonated (e.g., at low pH glycine is $^+\text{NH}_3\text{CH}_2\text{COOH}$) whereas at high pH both the amine and acid groups are deprotonated (e.g., $\text{NH}_2\text{CH}_2\text{COO}^-$). At intermediate pH amino acids exist in a zwitterionic form (*i.e.* a double ionic form, e.g., $^+\text{NH}_3\text{CH}_2\text{COO}^-$). In the space below sketch the titration curve for a solution of glycine titrated with NaOH, starting at low pH with the fully protonated form. Calculations are not required, however your sketch should demonstrate the successive equilibria which are involved in this titration. Write equations for the reactions which occur during the titration and indicate equivalence points and buffer region(s).

Feedback (Demonstrator to complete – Skills assessment – tick ALL that apply)**Core Skills**Analytical glassware: pipette - correct use (last chance)Scientific graphing: accurate line of best fit (or smooth curve if appropriate) appropriate scales correct title and axis labels (quantity/units) points accurately plottedTitration (last chance): setup (rinse, fill, no funnel or bubble) swirling during addition & pale endpoint**Non-core Skills**Applying chemical principles: correct logic and explanation used appropriate principle(s)Experimental accuracy: $\text{p}K_a$ within 0.5 of expected valueMastery (Experimental accuracy): $\text{p}K_a$ within 0.1 of expected valueProfessionalism: answers and discussion in pen locker and contents left cleanRecording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manualSafety awareness: safety eyewear always onTime management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

BUFFERS

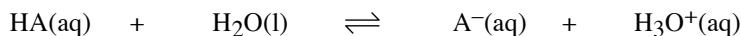
Aim

The aim of this experiment is to investigate the properties of buffer solutions and to learn about the methods for preparing buffer solutions.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

A **buffer solution** contains a weak acid and its conjugate base in approximately equal concentrations. The pH of a buffer solution does not change significantly when small quantities of H^+ or OH^- are added to it because the added H^+ or OH^- reacts with the base or acid respectively in the buffer. In any solution containing a conjugate acid/base pair (HA and A^-), this equilibrium occurs:



for which the equilibrium constant is $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$ (where stronger acids have larger K_a values)

which can be rewritten as: $\text{pH} = \text{p}K_a + \log_{10}\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$ where $\text{p}K_a = -\log_{10}(K_a)$. [equation 1]

An effective buffer solution needs to be able to consume added acid and base so the concentrations of the acid and conjugate base in the buffer need to be approximately equal: $[\text{A}^-]/[\text{HA}] \approx 1$ which means the pH of the buffer will be close to the $\text{p}K_a$ of the acid used to prepare it ($\log_{10}(1) = 0$ in equation 1).

When a small quantity of a strong acid is added to a buffer, it reacts with the A^- . For example, say you have a buffer made from a weak acid with $\text{p}K_a = 5.20$ and the concentrations of the weak acid and base are: $[\text{HA}] = 0.130 \text{ M}$ and $[\text{A}^-] = 0.130 \text{ M}$. The pH of this buffer will be 5.20 (from equation 1). If you add 1.0 mL of 0.10 M HCl (which contains 0.00010 mol H^+) to 100 mL of the buffer the following reaction happens:

	$\text{A}^-(\text{aq})$	$\text{H}_3\text{O}^+(\text{aq}, \text{added}) \rightarrow$	HA(aq)	$\text{H}_2\text{O(l)}$
Initial quantities	0.0130	0.00010	0.0130	mol
Final quantities	0.0129	0	0.0131	mol

Hence, the amount of HA increases slightly while that of A^- is decreased by the same amount. The H^+ from the strong acid is gone. The total volume of the solution is 101.0 mL and the concentrations are now: $[\text{HA}] = 0.0131 \text{ mol}/0.101 \text{ L} = 0.1297 \text{ M}$ and $[\text{A}^-] = 0.0129 \text{ mol}/0.101 \text{ L} = 0.1277 \text{ M}$ and the $\text{pH} = 5.20 + \log_{10}(0.1277/0.1297) = 5.19$. For comparison, if you add 1.0 mL of 0.10 M HCl to 100 mL of water, the pH of the solution changes from 7.0 to 3.0 (the actual concentration of H^+ increases by a factor of about 10,000 times).

Buffers can be prepared for almost any pH by appropriate choice of a weak acid and conjugate base pair. Since a buffer is most efficient at a pH close to the $\text{p}K_a$ of the acid (when the weak acid and conjugate base concentrations are approximately equal), this provides a way to select the appropriate pair for the desired pH of the buffer. The following table gives some common weak acid/base conjugate pairs and the pH range over which they can be used to prepare buffers.

Acid/Base Pair	$\text{p}K_a$ of Acid	pH Range of Effective Buffers
$\text{H}_3\text{PO}_4/\text{H}_2\text{PO}_4^-$	2.15	1.1 to 3.1
HF/F^-	3.17	2.2 to 4.2
HCOOH/HCO_2^-	3.74	2.7 to 4.7
$\text{CH}_3\text{COOH/CH}_3\text{CO}_2^-$	4.76	3.8 to 5.8
$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.20	5.9 to 7.9
$\text{NH}_4^+/\text{NH}_3$	9.24	8.2 to 10.2
$\text{HCO}_3^-/\text{CO}_3^{2-}$	10.33	9.3 to 11.3
$\text{HPO}_4^{2-}/\text{PO}_4^{3-}$	12.38	11.4 to 13.4

Buffers are most effective at pH values in the range $pK_a - 1$ to $pK_a + 1$, e.g., a buffer based on $\text{NH}_4^+/\text{NH}_3$ will only maintain a stable pH in the range 8.24 to 10.24. The further the pH of the buffer is away from the pK_a value, the less effective it is because the concentrations of weak acid and conjugate base must be unequal to produce such pH's.

Pre-lab Calculations (compulsory, must be completed to be allowed to start the experiment)

The tables below show the solutions you will have in the laboratory and the pH of the buffer you will need to prepare (based on your course and workspace). Select the reagents to prepare your buffer and in the space below calculate the volumes of each solution required to prepare 15.0 mL of the buffer. Circle the buffer pH and reagents you will use.

Solutions available to prepare buffers	
0.10 M CH_3COOH	0.10 M CH_3COONa
0.10 M NaH_2PO_4	0.10 M Na_2HPO_4
0.10 M NH_4Cl	0.10 M NH_3
0.10 M NaHCO_3	0.10 M Na_2CO_3

Workspace	Buffer pH (CHEM1011)	Buffer pH (CHEM1031/ CHEM1051)
1 and 6	4.76	4.50
2 and 7	7.20	7.30
3 and 8	9.24	9.40
4 and 9	10.33	10.40
5	4.76	4.60

Calculation of volumes required to make your buffer

Use equation 1, the pH of your buffer and the pK_a of the acid component of your buffer to calculate $[\text{A}^-]/[\text{HA}]$.

$$\text{pH} = pK_a + \log_{10}\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) \text{ which gives: } \log_{10}\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = \text{pH} - pK_a = \underline{\hspace{10cm}}$$

$$\text{which means that } \frac{[\text{A}^-]}{[\text{HA}]} = \underline{\hspace{10cm}}$$

Since the concentrations of the acids and conjugate bases are all the same in the solutions provided, the ratio $[\text{A}^-]/[\text{HA}]$ equals the ratio of the volume of base (V_b) to the volume of acid (V_a) (remember, A^- is the base).

$$[\text{A}^-]/[\text{HA}] = V_b/V_a = \underline{\hspace{2cm}} \text{ and to simplify the algebra we will call this ratio 'C'}$$

Volume of buffer = 15.0 mL, so $V_a + V_b = 15.0$ or $V_b = 15.0 - V_a$.

$V_b/V_a = C = (15.0 - V_a)/(V_a)$ so that $C \times V_a = (15.0 - V_a)$ which gives $V_a = 15.0 / (1 + C)$. Use this to calculate the volumes of acid and base solutions required to make 15.0 mL of your assigned buffer.

Volume of acid = mL Volume of base = mL

Calculation of acid and base concentrations in your buffer

Use the volumes above to calculate the concentration of acid and conjugate base in the buffer solution made by mixing the supplied acid and base solutions.

$$[\text{HA}] = 0.1 \text{ M} \times (V_a/15.0) = \underline{\hspace{10cm}}$$

$$[\text{A}^-] = 0.1 \text{ M} \times (V_b/15.0) = \underline{\hspace{10cm}}$$

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

The web page where you get your safety information has a link to a VIDEO about one of the techniques you will use in this experiment. **You should watch this video before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
0.10 M CH ₃ COOH		
0.10 M NaCH ₃ CO ₂		
0.10 M NaH ₂ PO ₄		
0.10 M Na ₂ HPO ₄		
0.10 M NH ₄ Cl		
0.10 M NH ₃		
0.10 M NaHCO ₃		
0.10 M Na ₂ CO ₃		
2 M HCl		
2 M NaOH		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Techniques

Please review the description of using the pH meters earlier in this manual (page 73).

Materials Needed

test tubes	0.10 M NaCH_3CO_2	0.10 M NH_4Cl	0.10 M CH_3COOH
10 mL measuring cylinder	0.10 M NaH_2PO_4	0.10 M NH_3	0.10 M Na_2CO_3
pH meter	0.10 M Na_2HPO_4	0.10 M NaHCO_3	2 M HCl 2 M NaOH



At the start of the experiment the pH electrode will be standing in a buffer solution of pH 4 in a conical flask. DO NOT CONTAMINATE THIS SOLUTION.

Do not throw out the solution. At the end of the experiment wash the electrode using a wash bottle and return it to this solution.



IMPORTANT: If you are using a Hanna 'stick' pH meter use narrow test tubes. If you are using a Mettler-Todeo pH meter use large (wide) test tubes.

Method

Part A: Preparation of Buffer

1. Using the results of your pre-lab calculation prepare 15.0 mL of the buffer solution assigned to your workspace in a large test tube and label it "Buffer solution A"
2. Transfer 5.0 mL of this buffer solution to a second test tube and add 5.0 mL of water to the tube. Label the tube "Buffer Solution B".
3. Place 10.0 mL of water in a third test tube. Label this tube "Water".
4. Repeat the above steps to prepare a second series of tubes containing Buffer Solution A, Buffer Solution B and water.
5. Complete the table on the results page, to show the composition of each buffer.

Part B: Investigation of Buffer Properties

1. Prepare 20 mL of 0.10 M HCl and 20 mL of 0.10 M NaOH in appropriately labelled large test tubes (using the 2 M HCl and NaOH solutions provided at your workspace). For some lab classes the 0.1 M HCl and 0.1 M NaOH may be provided – the lab supervisor will let you know if this is the case for your class.
2. Rinse the pH meter electrode and place it into one of the test tubes containing water. Record the pH of the water.

Addition of acid to water and Buffers A and B

3. Add 5 drops of the 0.10 M HCl to the water and record the pH. Repeat this for a further 5 drops of added HCl.
4. Rinse the electrode and transfer it to one of the tubes containing Buffer A. Record the pH of Buffer A.
5. Add 5 drops of the 0.10 M HCl to the buffer solution A and again record the pH.
6. Continue adding 0.10 M HCl to the tube of buffer A, until a total of 60 drops has been added, recording the pH at the volumes shown in the results table.
7. Repeat steps 5 to 7 for the first tube containing Buffer B .

Addition of base to water and Buffers A and B

8. Rinse the electrode, then repeat steps 3 to 8 using 0.10 M NaOH in place of the 0.10 M HCl, using the second set of tubes containing water, Buffer A, and Buffer B.
9. At the completion of the addition, remove the electrode from the solution, rinse thoroughly with water, and replace the electrode into the pH 7 buffer in the conical flask.
10. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.

RESULTS (Results must be written in pen.)

Your name: _____ Lab day/time: _____ Workspace: _____

Part A: Preparation of Buffer

Required pH of buffer	<input type="text"/>
Acid used	<input type="text"/>

Conjugate base used

Buffer Solution A

Concentration of acid	<input type="text"/>	Concentration of base	<input type="text"/>
-----------------------	----------------------	-----------------------	----------------------

(Note: these are the concentrations in the buffer, not the original solutions you used to make the buffer.)

Buffer Solution B

Concentration of acid	<input type="text"/>	Concentration of base	<input type="text"/>
-----------------------	----------------------	-----------------------	----------------------

Type of pH meter used (Toledo or Hanna):

Part B: Investigation of Buffer Properties

Addition of Acid (0.10 M HCl)

Water		Buffer Solution A		Buffer Solution B	
Total drops of 0.10 M HCl added	pH	Total drops of 0.10 M HCl added	pH	Total drops of 0.10 M HCl added	pH
0		0		0	
5		5		5	
10		10		10	
		20		20	
		30		30	
		40		40	
		50		50	
		60		60	

Addition of Base (0.10 M NaOH)

Water		Buffer Solution A		Buffer Solution B	
Total drops of 0.10 M NaOH added	pH	Total drops of 0.10 M NaOH added	pH	Total drops of 0.10 M NaOH added	pH
0		0		0	
5		5		5	
10		10		10	
		20		20	
		30		30	
		40		40	
		50		50	
		60		60	

Compare and explain the resistance of the buffer solutions to pH change with that of water.

Compare and explain the resistance to pH change of buffer solution A with that of buffer solution B.

Calculate the expected pH of your buffer solution A after 1.0 mL of 0.12 M HCl has been added to 15.0 mL of the buffer

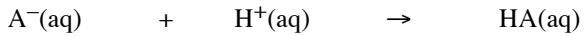
1. Calculate the amounts of acid and base in 15.0 mL of your buffer (use V_a and V_b from the pre-lab, converted to L):

Acid: $n_a = [HA] \times V_a = 0.10 \times V_a = \dots$ mol

Base: $n_b = [A^-] \times V_b = 0.10 \times V_b = \dots$ mol

2. Calculate the amount of acid added = $[H^+] \times V = 0.12 \text{ M} \times 0.0010 \text{ L} = \dots$ mol

3. Calculate the amounts of acid and conjugate base after all the **added** H^+ reacts:



Initial mol (from above)

Final 0 mol

4. Calculate the new pH:

Use equation 1 to calculate the final pH. You can take a shortcut because the acid and base are dissolved in the same volume of solution (16.0 mL), so $[HA] = \text{moles HA}/0.016 \text{ L}$ and $[A^-] = \text{moles } A^-/0.016 \text{ L}$. This means the volume cancels when you calculate the ratio $[A^-]/[HA]$ so you can use the final amounts of acid and base from above rather than their concentrations.

$$\text{pH} = pK_a + \log_{10}\left(\frac{[A^-]}{[HA]}\right) =$$

Buffers in Biochemistry [For CHEM1031 and CHEM1051 only]

A carbonic acid/hydrogen carbonate buffer is important in controlling the pH of blood. The normal pH of blood is 7.40. Calculate the ratio of carbonic acid to hydrogencarbonate at this pH given that the pK_a for the first dissociation of carbonic acid is 6.35.

Feedback (Demonstrator to complete – Skills assessment – tick ALL that apply)

Non-core Skills

Applying chemical principles: correct logic and explanation used appropriate principle(s)

Chemical calculations: correct arithmetic correct sig. fig. on results correct method

Mastery (Applying chemical principles): clear & concise explanation(s)

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

THERMOCHEMISTRY

Aim

The aim of this set of experiments is to use calorimetry to measure the enthalpy change for a number of chemical reactions and then to verify that these enthalpy changes obey Hess's Law.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

Most chemical reactions either liberate heat as they occur (exothermic reactions) or they absorb heat (endothermic reactions). The amount of heat liberated or absorbed by a particular reaction can be determined by **calorimetry**. A **calorimeter** is a thermally insulated container which does not exchange heat with its surroundings. Consider an exothermic reaction carried out in a calorimeter. As the reaction proceeds, heat is released. As the heat cannot escape from the calorimeter, it goes into warming the contents of the calorimeter and the calorimeter itself, leading to a rise in temperature of the calorimeter and its contents, which can be measured using a thermometer. The heat necessary to raise the temperature of the calorimeter and its contents can be calculated from the equation:

$$q = m \times C_{\text{soln}} \times \Delta T + C_{\text{cal}} \times \Delta T$$

where m = mass of the solution, C_{soln} = heat capacity of the solution, ΔT = temperature change and C_{cal} = calorimeter heat capacity, also known as the **calorimeter constant**. The calorimeter constant is specific for a particular calorimeter and is a measure of the heat absorbed by the calorimeter in order to change its temperature by 1 K. As energy is conserved, the heat which goes into warming the calorimeter and its contents must equal the heat *released* by the reaction ($-q_{\text{reaction}}$). Thus:

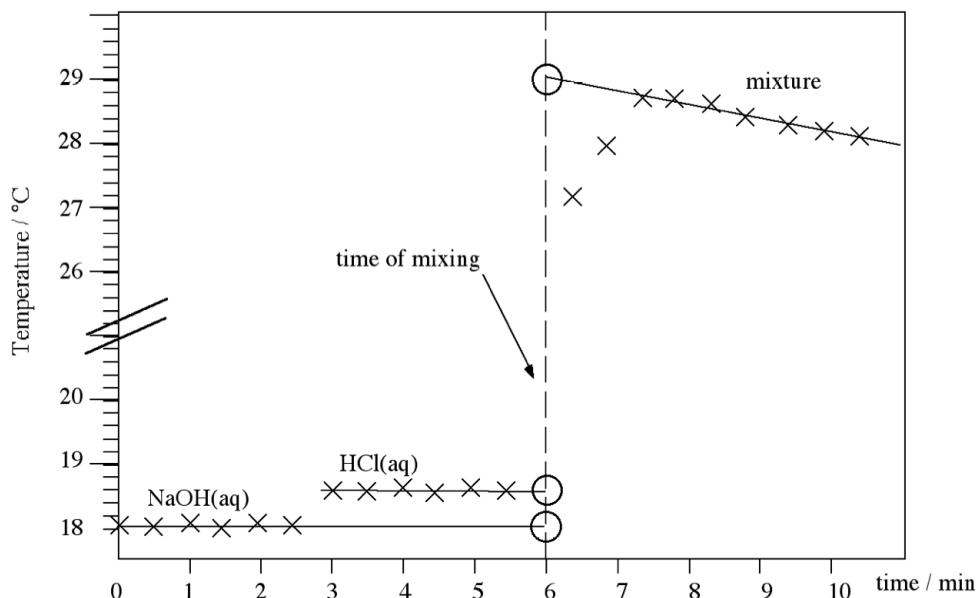
$$q_{\text{reaction}} = -(m \times C_{\text{soln}} \times \Delta T + C_{\text{cal}} \times \Delta T)$$

Since the reaction is carried out at constant pressure (because the calorimeter is open to the air), by definition the heat equals the enthalpy change for the reaction, symbolized by ΔH . Thus:

$$\Delta H = -(m \times C_{\text{soln}} \times \Delta T + C_{\text{cal}} \times \Delta T)$$

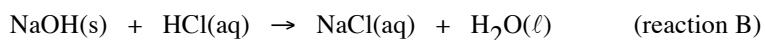
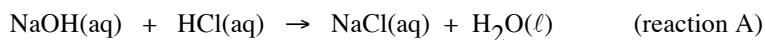
One practical problem with calorimetry lies in the determination of the temperature change. The reaction being studied may be slow and it may take some time for the system to reach its maximum temperature. The thermometer reading is likely to lag behind the actual solution temperature and there will be some heat loss from the calorimeter. These problems can be overcome by measuring the temperature of the reactants and products periodically and then extrapolating the temperatures graphically to the time of mixing. 'Extrapolate' means to estimate the value a quantity has outside the times it was measured. The graph below shows how this can be done.

Temperature versus time for NaOH(aq) and HCl(aq) reacting.



Each 'x' indicates a measured temperature for a solution. By drawing a straight line through the x's and extending (extrapolating) that line to the time the solutions were mixed, the temperatures of the solutions at the time of mixing can read from the graph (at the circled points). Note also that the axes are labeled with the quantity and units, and the scale on the temperature axis is broken to allow the temperatures to be plotted to ± 0.2 °C, without requiring an enormous piece of graph paper. Because you have several lines on your graph, the data from this experiment are best plotted by hand. It is difficult to use a program to process this data.

Hess's law states that the enthalpy change for a chemical reaction equals the sum of the enthalpy changes for any series of reactions which can be combined to give the reaction of interest. So if you carry out a reaction as a series of steps and add up the enthalpy changes for the steps, you will get the same enthalpy change as you would by carrying out the reaction as a single step. Here are the reactions you will be measuring enthalpy changes for in this experiment. You should be able to see that two of these reactions can be combined to give the other one.



Calorimetry is used in many disciplines and industries. The calorific values of foods are determined by burning food in a closed calorimeter, as are the calorific values of fuels such as coal. In medicinal chemistry calorimetry is used to determine energy changes when drug candidates bind to their target sites, providing information about the type of interaction between the drug and its target.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
1 M HCl		
1 M NaOH		
solid NaOH		
litmus paper		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
--	--

Materials Needed

standardized sodium hydroxide (approx. 1 M)	calorimeter	digital thermometer ($\pm 0.1^\circ\text{C}$)
standardized hydrochloric acid (approx. 1 M)	250 mL beakers	watch (etc.) with a precision of ± 10 s
solid NaOH	100 mL measuring cylinder	(or use clock provided on lab screens)
red and blue litmus paper	watch glass	glass stirring rod

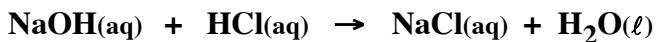


The key to getting good results in this experiment is to stir the solutions thoroughly and continuously while making your temperature measurements.

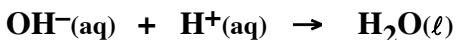
You can do parts A, B, and C in any order, to avoid queuing for solutions or balances.

Method

Part A – ΔH for Hydrochloric Acid reacting with Aqueous Sodium Hydroxide.



or



1. Clean and dry the calorimeter.
2. Dispense 60 mL of the standardized sodium hydroxide solution directly into the dry, clean calorimeter.
3. Dispense 50 mL of the standardized hydrochloric acid directly into a clean, dry beaker. Record in pen the exact concentration of the acid used onto your results page.



Record the exact concentrations of the hydrochloric acid and sodium hydroxide used.

4. Measure and record to 1 decimal place the temperature of the sodium hydroxide solution at 0.5 minute intervals for a total of 6 measurements. Wash and dry the thermometer. Next, measure and record the temperature of the hydrochloric acid solution at 0.5 minute intervals for a total of 6 measurements.

5. At time = 6.0 minutes, add the hydrochloric acid solution to the sodium hydroxide solution in the calorimeter. Mix well using a *glass stirring rod* and stir the solution for 15 seconds.

Caution: Do not use the thermometer to stir the solution.

6. At time = 6.5 minutes, (and at further 0.5 minute intervals), measure the temperature of the solution. Stir the solution using the stirring rod between the temperature measurements. Continue to measure the temperature until a steady decrease in temperature is noted or until time = 11.0 minutes.

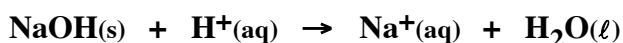
Note: In order to obtain a smooth temperature curve you must **stir the solution continuously** between temperature measurements and keep the thermometer in the solution the whole time.

7. Rinse (under the tap) the temperature probe on the thermometer and dry it.
8. Obtain a piece of blue litmus paper and a piece of red litmus paper. Transfer a drop of the final solution onto both pieces of litmus paper. Use the colours to determine whether the resultant solution is acidic or basic and record this in the results.
9. Graph the results on the results page. Be sure to allow sufficient range on the temperature axis so that the temperature at the time of mixing fits on the graph.
10. Extrapolate your lines to the time of mixing (time = 6.0 minutes) and determine the change in temperature as shown on the graph in the introduction.
11. Calculate the enthalpy change for the reaction by completing the calculations in the results pages.

Part B – ΔH for Hydrochloric Acid reacting with solid Sodium Hydroxide.



or



Solid NaOH is very corrosive

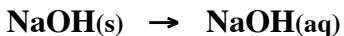


Exercise care in handling solid NaOH. It is corrosive to equipment such as balances and will cause burns on contact with skin.

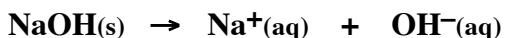
All spillage must be cleaned up immediately and if skin contact occurs wash immediately with water.

1. Using the top loading balance which reads to 2 decimal places, weigh a clean sample tube and cap. DO NOT TARE THE BALANCE, you need the actual mass of the sample tube plus NaOH because you will re-weigh the emptied tube in step 6. Weigh anywhere between 1.9 – 2.1 g of solid sodium hydroxide pellets into this sample tube and cap immediately. Record the masses in pen on your results page.
 2. Dispense 100 mL of the standardized hydrochloric acid solution directly into the clean dry calorimeter. Measure and record to 1 decimal place the temperature of the standardized hydrochloric acid solution at 0.5 minute intervals.
 3. At elapsed time = 3.5 minutes, add the solid sodium hydroxide pellets to the hydrochloric acid solution in the calorimeter. Start stirring the mixture with a glass stirring rod. Keep the thermometer in the solution and **do not stop stirring** the mixture until all the temperature readings have been obtained.
- Caution:** Do not use the thermometer to stir the solution.
4. At elapsed time = 4.0 minutes, measure the temperature of the resulting solution. Record the temperature and repeat these temperature measurements at 0.5 minute intervals until either a steady decline in temperature is noted or until 10.0 minutes has elapsed.
 5. Rinse (under the tap) the temperature probe on the thermometer and dry it.
 6. Obtain a piece of blue litmus paper and a piece of red litmus paper. Transfer a drop of the final solution onto both pieces of litmus paper. Use the colours to determine whether the resultant solution is acidic or basic and record this in the results.
 7. Re-weigh the empty sample tube from step 3; (with the cap). Record the mass on the results page.
 8. Extrapolate your lines to the time of mixing (time = 3.5 minutes) and determine the change in temperature as shown on the graph in the introduction. Be sure to allow sufficient range on the temperature axis so that the temperature at the time of mixing fits on the graph.
 9. Calculate the enthalpy change for the reaction by completing the calculations in the results pages.

Part C – ΔH for dissolution of solid Sodium Hydroxide.



or



Solid NaOH is very corrosive



Exercise care in handling solid NaOH. It is corrosive to equipment such as balances and will cause burns on contact with skin.

All spillage must be cleaned up immediately and if skin contact occurs wash immediately with water.

1. Using the top loading balance which reads to 2 decimal places, weigh a clean sample tube and cap. DO NOT TARE THE BALANCE, you need the actual mass of the sample tube plus NaOH because you will re-weigh the emptied tube in step 6. Weigh anywhere between 1.9 – 2.1 g of solid sodium hydroxide pellets into this sample tube and cap immediately. Record in pen the masses on your results page.
2. Measure 100.0 mL of water into the clean dry calorimeter.
3. Measure and record to 1 decimal place the temperature of the water at 0.5 minute intervals.
4. At time = 3.5 minutes, add the solid sodium hydroxide to the water in the calorimeter. Start stirring the mixture with a glass stirring rod. Keep the thermometer in the solution and **do not stop stirring** the mixture until all the temperature readings have been obtained.
Caution: Do not use the thermometer to stir the solution.
5. At time = 4.0 minutes, and at subsequent 0.5 minute intervals, measure and record the temperature of the resulting solution. Between temperature measurements stir the solution thoroughly.
6. Re-weigh the empty sample tube (with the cap) from step 4. Record the mass on your results page.
7. Rinse (under the tap) the temperature probe on the thermometer and dry it.
8. Extrapolate your lines to the time of mixing (time = 3.5 minutes) and determine the change in temperature as shown on the graph in the introduction. Be sure to allow sufficient range on the temperature axis so that the temperature at the time of mixing fits on the graph.
9. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.
10. Determine the enthalpy of solution by completing the calculation in the results section.

Blank page.

(Can be used for working or calculations)

School of Chemistry, UNSW

Coversheet for Submission of Individual Report

Report: Thermochemistry

Circle your course: CHEM1011 CHEM1031 CHEM1051

Student Name (full name): _____

Student Number (e.g., z1234567): _____ Lab. class: day/time: _____

Demonstrator: _____ Workspace: _____

Date and Time Submitted: _____

In preparing this assessment task I have followed the Student Code of Conduct. I certify that I have read and understand the University requirements in respect of student academic misconduct outlined in the Student Code of Conduct and Annexure 1 of the Student Misconduct Procedures. I declare that this assessment item is my own work, except where acknowledged, and has not been submitted for academic credit previously in whole or in part. I acknowledge that the assessor of this item may, for assessment purposes: (1) provide a copy to another staff member of the University; (2) communicate a copy of this assessment item to a plagiarism checking service (such as Turnitin) which may then retain a copy of the assessment item on its database for the purpose of future plagiarism checking. I have retained a copy of this, my assignment, which I can provide if necessary. By signing this declaration I am agreeing to the statements and conditions above.

Student signature and date: _____

Full UNSW policy concerning student misconduct is available at: <<https://student.unsw.edu.au/conduct>>

Feedback (Demonstrator to complete – Skills assessment – tick ALL that apply)

Core Skills

Scientific graphing (last chance): accurate line of best fit (or smooth curve if appropriate) appropriate scales
 correct title and axis labels (quantity/units) points accurately plotted

Non-core Skills

Applying chemical principles: correct logic and explanation

Chemical calculations: correct arithmetic correct method correct sig. fig. on results correct units

Experimental accuracy: acceptable agreement with Hess's Law

Mastery (Applying chemical principles): clear & concise explanation(s)

Mastery (Experimental accuracy): accurate agreement with Hess's Law

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

RESULTS (Results must be written in pen.)

PART A – ΔH for Hydrochloric Acid reacting with Aqueous Sodium Hydroxide.

Concentration of HCl(aq): _____

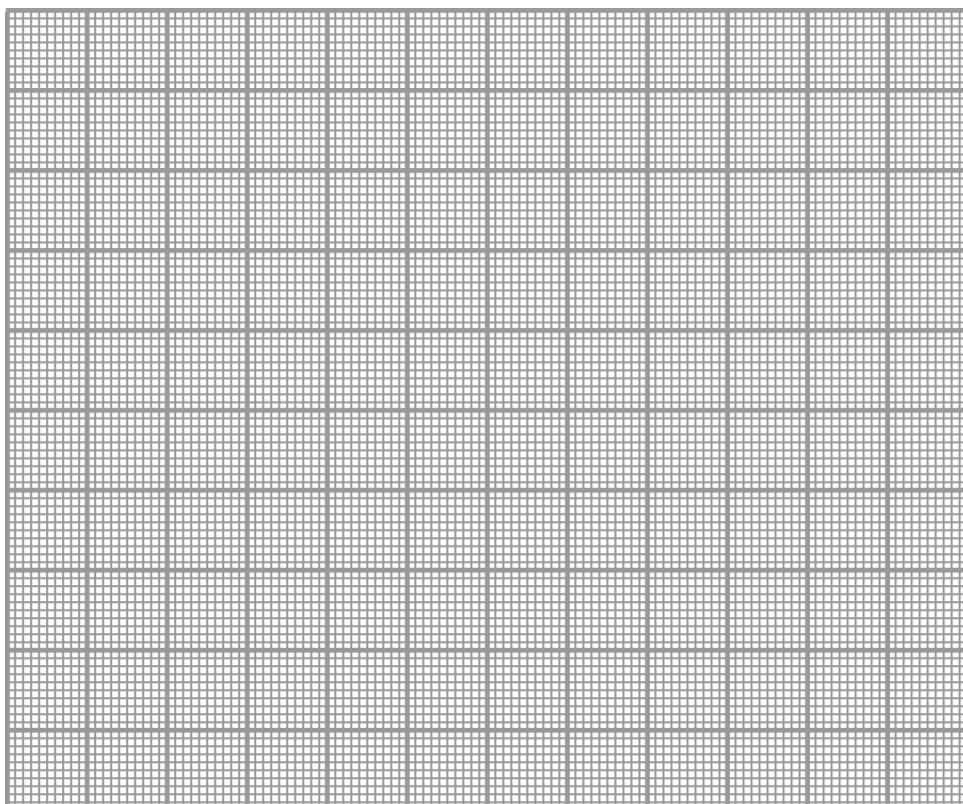
Concentration of NaOH(aq): _____

Elapsed time / min	0	0.5	1.0	1.5	2.0	2.5
Temperature of aqueous NaOH / °C						

Elapsed time / min	3.0	3.5	4.0	4.5	5.0	5.5
Temperature of HCl / °C						

Elapsed time at mixing = 6.0 min

Elapsed time / min	6.5	7.0	7.5	8.0	8.5	9.0	9.5	10.0	10.5	11.0
Temp. of mixture / °C										



Colour of litmus with final solution: _____

Is final solution acidic or basic? _____

Temperatures/°C of solutions obtained from the graph, extrapolated to time of mixing (as shown in the introduction):

NaOH(aq):_____ HCl(aq):_____ mixture:_____

Calculate ΔT from the temperatures extrapolated to the time of mixing: ΔT = temperature of mixture – average of temperatures of NaOH(aq) and HCl(aq) solutions

$$= \underline{\hspace{10cm}}$$

Calculate the number of moles of H^+ neutralised:

$$[HCl] / \text{mol L}^{-1} = \underline{\hspace{2cm}} \quad \text{Volume of HCl solution/mL} = \underline{\hspace{2cm}}$$

$$n_{\text{HCl}} / \text{mol} = \underline{\hspace{2cm}}$$

$$[\text{NaOH}] / \text{mol L}^{-1} = \underline{\hspace{2cm}} \quad \text{Volume of NaOH solution/mL} = \underline{\hspace{2cm}}$$

$$n_{\text{NaOH}} / \text{mol} = \underline{\hspace{2cm}}$$

$$\therefore \text{Limiting reagent is } \underline{\hspace{2cm}} \quad \text{amount } H^+ \text{ reacted / mol} = \underline{\hspace{2cm}}$$

Knowing which reagent is in excess, should the final solution be acidic or basic? _____

Is this consistent with your observations and why? _____

Calculation of Enthalpy of Neutralization

Note: include units and a justifiable number of significant figures for all quantities you calculate. Keep intermediate results in your calculator to avoid roundoff error.

1. Use the total volume of solution and its density to calculate the mass of the final solution:

$$\text{Volume of final solution/mL} = \underline{\hspace{2cm}} \quad \text{Density} = 1.02 \text{ g mL}^{-1} \text{ at } 20^\circ\text{C}$$

$$\text{Mass of final solution} = \underline{\hspace{2cm}}$$

2. Use the mass of the final solution, ΔT and the heat capacity of the solution to calculate the heat absorbed by the solution.

$$\text{Mass of final solution /g} = \underline{\hspace{2cm}} \quad C_p = 4.04 \text{ J g}^{-1} \text{ K}^{-1}$$

$$\text{Heat absorbed by solution} = \underline{\hspace{2cm}}$$

3. Use the calorimeter constant and ΔT to calculate the heat absorbed by the calorimeter.

$$\text{Calorimeter constant} = 30 \text{ J K}^{-1}$$

$$\text{Heat absorbed by calorimeter} = \underline{\hspace{2cm}}$$

4. Calculate the heat associated with the chemical reaction, which equals the negative value of the sum of the heats calculated in the two previous steps, that is $-(\text{heat absorbed by the calorimeter} + \text{heat absorbed by the solution})$.

$$\text{Heat associated with reaction} = \underline{\hspace{2cm}}$$

5. The enthalpy change equals the heat calculated in the previous step (by definition enthalpy change equals the heat associated with a reaction when the reaction is carried out at constant pressure, as you did in this experiment).

$$\text{Enthalpy change for the reaction in the calorimeter} = \underline{\hspace{2cm}}$$

6. The molar enthalpy change is the enthalpy change for the reaction which occurred in the calorimeter divided by the amount (mol) of H^+ reacted.

$$\text{Molar enthalpy change for the reaction} = \underline{\hspace{2cm}}$$

PART B – ΔH for Hydrochloric Acid reacting with solid Sodium Hydroxide.

Mass of empty sample tube /g _____

Concentration of HCl(aq): _____

Mass of sample tube + NaOH(s) /g _____

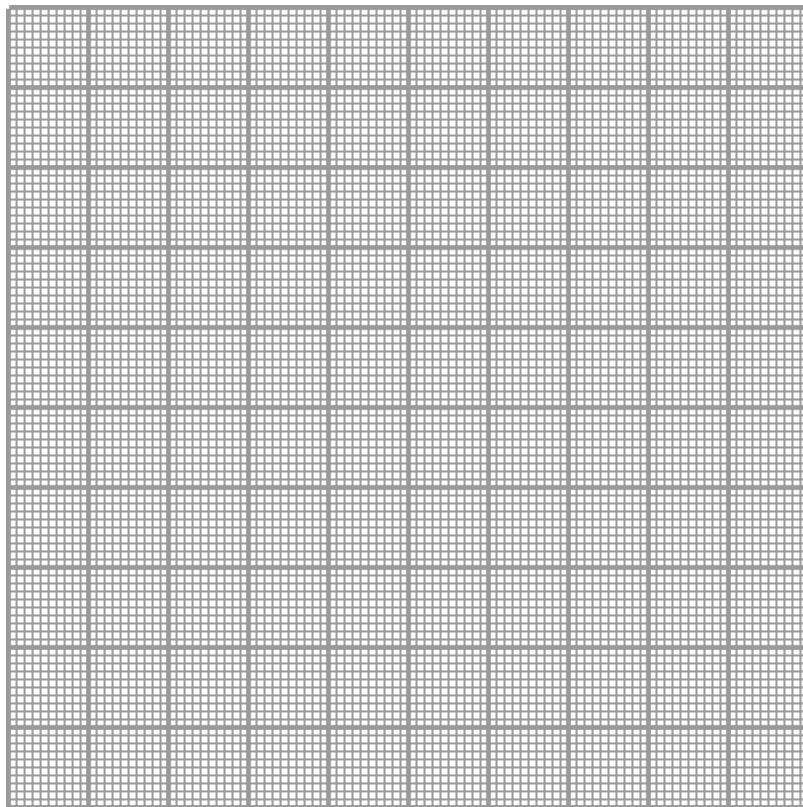
Mass of sample tube after emptying NaOH /g _____

Mass of NaOH(s) added to calorimeter/g _____

Elapsed time / min	0	0.5	1.0	1.5	2.0	2.5	3.0
Temperature of hydrochloric acid / °C							

Elapsed time at mixing = 3.5 min

Elapsed time / min	4.0	4.5	5.0	5.5	6.0	6.5	7.0	7.5	8.0	8.5	9.0	9.5	10.0
Temp. of mixture / °C													



Colour of litmus with final solution: _____ Is final solution acidic or basic? _____

Temperatures/°C of solutions obtained from the graph, extrapolated to time of mixing (as shown in the introduction):

HCl(aq): _____ mixture: _____

Calculate ΔT from the temperatures extrapolated to the time of mixing:

ΔT = temperature of mixture – temperature of HCl(aq)

$$= \underline{\hspace{10cm}}$$

Calculate the number of moles of H⁺ neutralised:

$$[\text{HCl}] / \text{mol L}^{-1} = \underline{\hspace{2cm}} \quad \text{Volume of HCl solution/mL} = \underline{\hspace{2cm}}$$

$$n_{\text{HCl}} / \text{mol} = \underline{\hspace{2cm}}$$

$$\text{mass NaOH / g} = \underline{\hspace{2cm}}$$

$$n_{\text{NaOH}} / \text{mol} = \underline{\hspace{2cm}}$$

$$\therefore \text{Limiting reagent is } \underline{\hspace{2cm}} \quad \text{amount H}^+ \text{ reacted / mol} = \underline{\hspace{2cm}}$$

Knowing which reagent is in excess, should the final solution be acidic or basic? _____

Is this consistent with your observations and why? _____

Calculation of Enthalpy of Neutralization

Note: include units and a justifiable number of significant figures for all quantities you calculate. Keep intermediate results in your calculator to avoid roundoff error.

1. Calculate the mass of the final solution:

$$\text{Volume of HCl(aq)/mL} = \underline{\hspace{2cm}} \quad \text{Density} = 1.02 \text{ g mL}^{-1} \text{ at } 20^\circ\text{C}$$

$$\text{Mass of HCl(aq) /g} = \underline{\hspace{2cm}}$$

$$\text{Mass of NaOH dissolved /g} = \underline{\hspace{2cm}}$$

$$\text{Mass of final solution} = \text{mass of HCl(aq)} + \text{mass of NaOH(s)} = \underline{\hspace{2cm}}$$

2. Calculate the heat absorbed by the solution.

$$\text{Mass of final solution/g} = \underline{\hspace{2cm}} \quad C_p = 4.01 \text{ J g}^{-1} \text{ K}^{-1}$$

$$\text{Heat absorbed by solution} = \underline{\hspace{2cm}}$$

3. Calculate the heat absorbed by the calorimeter.

$$\text{Calorimeter constant} = 30 \text{ J K}^{-1}$$

$$\text{Heat absorbed by calorimeter} = \underline{\hspace{2cm}}$$

4. Calculate the heat associated with the chemical reaction, which equals the negative value of the sum of the heats calculated in the two previous steps, that is –(heat absorbed by the calorimeter + heat absorbed by the solution).

$$\text{Heat associated with reaction} = \underline{\hspace{2cm}}$$

5. The enthalpy change equals the heat calculated in the previous step (by definition enthalpy change equals the heat associated with a reaction when the reaction is carried out at constant pressure, as you did in this experiment).

$$\text{Enthalpy change for the reaction in the calorimeter} = \underline{\hspace{2cm}}$$

6. The molar enthalpy change is the enthalpy change for the reaction which occurred in the calorimeter divided by the amount (mol) of H⁺ reacted.

$$\text{Molar enthalpy change for the reaction} = \underline{\hspace{2cm}}$$

PART C – ΔH for dissolution of solid Sodium Hydroxide.

Mass of sample tube /g _____

Mass of sample tube + NaOH(s) /g _____

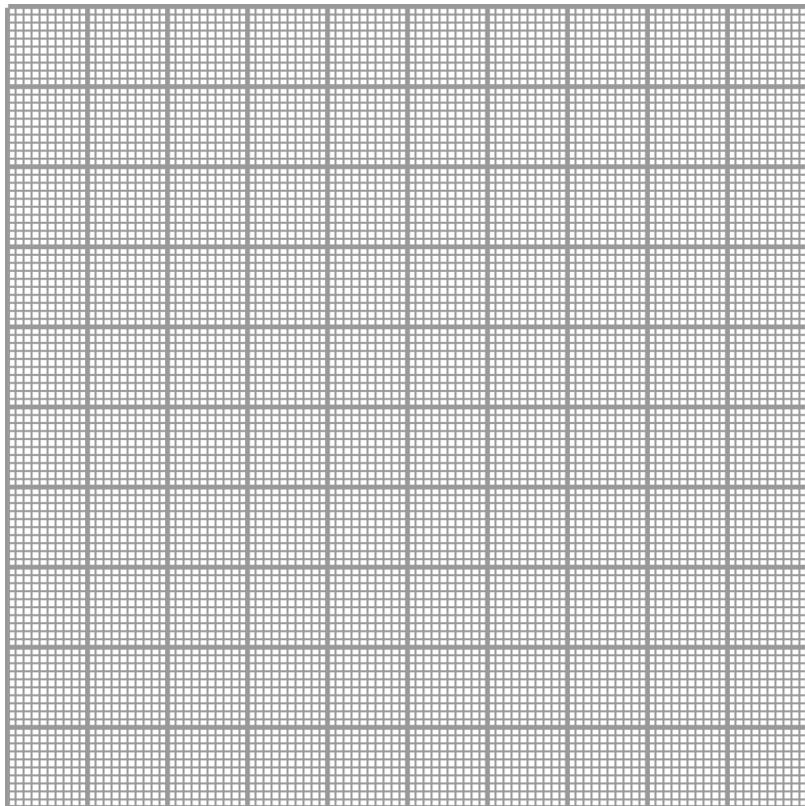
Mass of sample tube after emptying NaOH /g _____

Mass of NaOH(s) added to calorimeter/g _____

Elapsed time / min	0	0.5	1.0	1.5	2.0	2.5	3.0
Temperature of water / °C							

Elapsed time at mixing = 3.5 min

Elapsed time / min	4.0	4.5	5.0	5.5	6.0	6.5	7.0	7.5	8.0	8.5	9.0	9.5	10.0
Temp. of mixture / °C													



Temperatures/°C of solutions obtained from the graph, extrapolated to time of mixing (as shown in the introduction):

H₂O(l): _____ NaOH(aq): _____

Calculate ΔT from the temperatures extrapolated to the time of mixing:

ΔT = temperature of NaOH(aq) – temperature of H₂O(ℓ)

$$= \underline{\hspace{10cm}}$$

Calculate the number of moles of NaOH dissolved:

mass NaOH / g = _____

$$n_{\text{NaOH}} / \text{mol} = \underline{\hspace{10cm}}$$

Calculation of Enthalpy of Reaction

Note: include units and a justifiable number of significant figures for all quantities you calculate. Keep intermediate results in your calculator to avoid roundoff error.

1. Calculate the mass of the final solution:

Volume of water / mL = _____ Density = 0.997 g mL⁻¹

Mass of water / g = _____

Mass of NaOH dissolved / g = _____

Mass of final solution = mass of water + mass of NaOH(s) = _____

2. Calculate the heat absorbed by the solution.

Mass of final solution / g = _____ $C_p = 4.07 \text{ J g}^{-1} \text{ K}^{-1}$

Heat absorbed by solution = _____

3. Calculate the heat absorbed by the calorimeter.

Calorimeter constant = 30 J K⁻¹

Heat absorbed by calorimeter = _____

4. Calculate the heat associated with the chemical reaction, which equals the negative value of the sum of the heats calculated in the two previous steps, that is –(heat absorbed by the calorimeter + heat absorbed by the solution).

Heat associated with reaction = _____

5. The enthalpy change equals the heat calculated in the previous step (by definition enthalpy change equals the heat associated with a reaction when the reaction is carried out at constant pressure, as you did in this experiment).

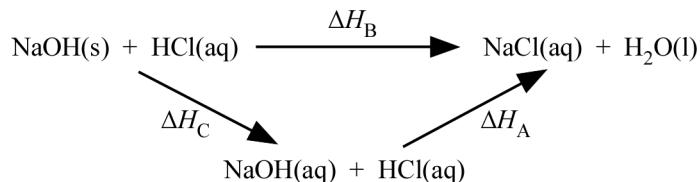
Enthalpy change for the reaction in the calorimeter = _____

6. Calculate the molar enthalpy change.

Molar enthalpy change for the dissolution of NaOH(s) = _____

Application of Hess's Law

The diagram below shows the three processes studied in this experiment. The letters 'A', 'B', and 'C' refer to the parts of the experiment.



Use Hess's Law to write (in the space below) an algebraic expression for ΔH_B in terms of ΔH_A and ΔH_C .

According to Hess's Law:

From your experiment: $\Delta H_A = \dots \text{ kJ mol}^{-1}$ $\Delta H_B = \dots \text{ kJ mol}^{-1}$ $\Delta H_C = \dots \text{ kJ mol}^{-1}$

By substituting your experimental values into this relationship, determine if (within experimental uncertainty) your values obey Hess's Law. If the largest source of uncertainty is in ΔT (because of the manual extrapolation), an uncertainty of 0.1 °C in ΔT produces an uncertainty of approximately 1 kJ mol⁻¹ in ΔH in this experiment. If you believe your ΔT values have a larger uncertainty you can use this to estimate the uncertainties in each of your ΔH values.

[For CHEM1031 and CHEM1051, optional for CHEM1011]

What effect, if any, would using nitric acid instead of hydrochloric acid have on the enthalpy changes measured in this experiment? Include some justification for your answer.

GALVANIC CELLS

Aim

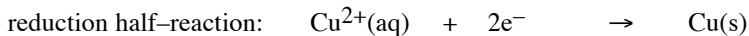
In this experiment you will construct several galvanic cells in order to learn about:

- the physical construction of galvanic cells and the correspondence with standard notation for cells,
- using the measured value of the cell potentials for a series of metal ion|metal half cells to calculate the standard reduction potentials of the metals, and
- (CHEM1031/CHEM1051 only) using a measured cell potential to determine an equilibrium constant.

See the Feedback panel in the Results section for a complete list of skills being assessed in this experiment.

Introduction

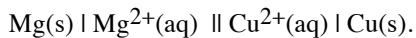
What are oxidation and reduction, and electrochemical cells? Oxidation–reduction (redox) reactions involve the transfer of electrons between substances. The substance which **loses** electrons is **oxidized** and the substance which **gains** those electrons is **reduced**. If the oxidation and reduction processes occur in different places, the electrons will have to flow through an external circuit to get from the site of the oxidation to the site of the reduction. This happens in a electrochemical cell. If an electrochemical cell generates a current it is called a **galvanic cell**. If an electrochemical cell requires an external source of electrical energy to drive the chemical reaction it is an **electrolytic cell**. For example, if you add powdered magnesium to a solution containing Cu^{2+} ions, the magnesium will lose electrons (be oxidized) and the Cu^{2+} will gain those electrons (be reduced) and heat will be generated. The overall reaction can be written as a combination of two half reactions:



If you bring together the substances involved in each half reaction in separate beakers (Mg dipping into a solution of Mg^{2+} in one beaker, and Cu dipping into a solution of Cu^{2+} in another beaker), and connect the metals by a wire, and connect the beakers by a strip of paper dipped in an electrolyte (e.g., sodium sulfate) then a current will flow through the wire and less heat will be generated by the reaction. The conducting paper strip (**a salt bridge**) is needed to avoid charge building up in the beakers.

Cell notation. To describe a cell you could write the chemical equation for the reaction occurring in the cell, however there are aspects of the construction of the cell which can affect the voltage it produces, so a more detailed notation is used. In **standard cell notation**, the substances in the cell are listed starting at the electrode where oxidation occurs, working through the cell to the electrode where reduction occurs, with each phase boundary denoted by a vertical line (hence this notation is sometimes called 'line' notation).

For a cell in which magnesium metal is oxidized to Mg^{2+} and Cu^{2+} is reduced to copper metal, the cell diagram would be:



The solutions of Mg^{2+} and Cu^{2+} are in separate containers joined by a salt bridge, which is represented by a double vertical line. In electrochemistry the **anode** is defined as the electrode where oxidation occurs, and the **cathode** is where reduction occurs. In this cell the $\text{Mg}^{2+}(\text{aq}) | \text{Mg(s)}$ half cell is the anode and the $\text{Cu}^{2+}(\text{aq}) | \text{Cu(s)}$ half cell is the cathode.

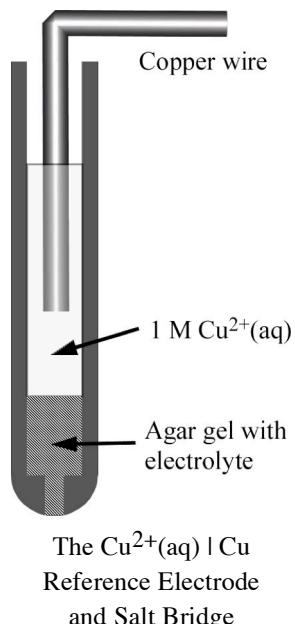
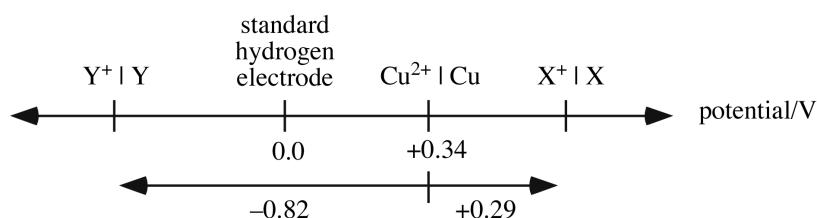
The cell potential is given by the **potential of the right-hand electrode minus the potential of the left-hand electrode**. If the right-hand electrode is the more positive of the two electrodes, the cell potential is a positive quantity and this corresponds to a spontaneous reaction with oxidation happening in the left side half cell and reduction happening in the right side half cell.

In part A of this experiment you will construct a simple galvanic cell, measure the potential difference it produces and verify the nature of the spontaneous chemical reaction which is occurring in the cell. You will be able to see the correspondence between the physical construction of a galvanic cell and the standard notation for cells.

Part B of this experiment illustrates the relationship between standard half-cell potentials and standard cell potentials. Relative to the standard hydrogen cell (which by definition has a standard half-cell potential of zero volts) the $\text{Cu}^{2+}|\text{Cu}$ half cell has a potential of 0.34 V. By connecting other half cells to the copper half cell and measuring the overall cell voltage it is possible to determine the half cell potentials for any redox couple.

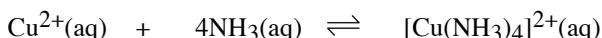
The copper reference electrode consists of a narrow glass tube constricted at one end and plugged with agar gel containing an electrolyte, sodium nitrate. A small piece of cotton wool gives strength to the gel. Above the gel is 1M CuSO₄ solution into which is inserted copper wire which extends out of the top of the tube to form the terminal for the half cell. This assembly is a reference electrode with a built-in salt bridge which can be placed into the electrolyte of another half cell where the agar–NaNO₃ bridge connects the two half cells without allowing them to come in contact (because the solutions do not penetrate through the gel). The electrolyte (NaNO₃) in the gel allows the balance of positive and negative ions in each half-cell to be maintained when the cell is in operation.

The standard reduction potential for a half cell is defined as the potential of the half cell with respect to the standard hydrogen electrode. A standard hydrogen electrode is not easy to setup, so in this experiment the standard copper electrode described above is used. This means that you need to correct the potential differences you will measure when you calculate the standard reduction potentials of the metals connected to the copper electrode. One way to visualize this is to think of the potentials of these electrodes displayed on a number line with the standard hydrogen electrode at zero.



For example, say a cell made up of the Cu²⁺|Cu and X⁺|X half cells generates 0.29 V and the X⁺|X half-cell is positive with respect to the copper electrode. This places the X⁺|X half cell 0.29 V to the right of the copper electrode on the above diagram. If a cell made up of Y⁺|Y and Cu²⁺|Cu generates 0.82 V and the copper electrode is positive with respect to the Y⁺|Y half-cell then Y⁺|Y is -0.82 V relative to the copper electrode (0.82 V to the left of the copper electrode). In order to calculate the standard reduction potentials of these two half cells you need to add the potential of the copper half cell with respect to the hydrogen electrode onto the measured cell potentials. Thus the standard reduction potential for X⁺ | X is +0.63 V and that for Y⁺ | Y is -0.48 V. It is crucial to determine not only the size of the cell potential, but also the polarity of the electrode (X or Y) with respect to the copper electrode.

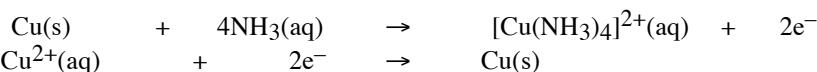
In part C of the experiment (**for CHEM1031 and CHEM1051 only**), measurement of a cell potential will be used to determine the value of the equilibrium constant for the following reaction:



The equilibrium constant for this type of reaction is called a stability constant:

$$K_{stab} = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}}{[\text{Cu}^{2+}][\text{NH}_3]^4} \text{ where all the concentrations are measured at equilibrium.}$$

The reaction can be separated into oxidation and reduction half reactions:

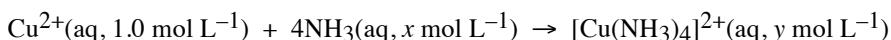


These two half reactions take place in this cell:



where the cell voltage *E* can be measured. If all the concentrations were at their equilibrium values, then *E* would be zero. Because this copper complex is quite stable this would mean the concentration of Cu²⁺(aq) would be very small. When you setup this cell you will use the copper reference electrode described above, so that [Cu²⁺(aq)] = 1.0 mol L⁻¹. The concentrations of [Cu(NH₃)₄]²⁺(aq) and NH₃(aq) can be calculated from the initial concentrations of Cu²⁺(aq) and NH₃ mixed together to form the complex, with Cu²⁺ being the limiting reagent and assuming the reaction goes to completion.

The reaction happening in the cell, with concentrations specified, is:



where the concentrations of ammonia (*x*) and the complex (*y*) are those in the left-hand half-cell, calculated from the stoichiometry of the formation of the complex.

Using the Nernst equation: $E = E^\circ - \frac{RT}{nF} \ln Q$ combined with $RT \ln K = nFE^\circ$

gives: $E = \frac{RT}{nF} \ln K - \frac{RT}{nF} \ln Q$

which rearranges to: $\ln K = \frac{nFE}{RT} + \ln Q$

This shows that from a measurement of the cell voltage the value of the equilibrium constant can be calculated. In this expression the reaction quotient, Q , has the same form as the equilibrium constant:

$$Q = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}}{[\text{Cu}^{2+}][\text{NH}_3]^4}$$

where the concentrations are those in your cell (*which is not at equilibrium*).

The concentration of Cu^{2+} is that in the reference cell, while the concentrations of NH_3 and $[\text{Cu}(\text{NH}_3)_4]^{2+}$ are those in the other half-cell, calculated as outlined above.

Safety Information (compulsory, must be completed to be allowed to start the experiment)

Before starting this experiment, you must use the pre-lab web page (link in Moodle) to look up the precautions associated with the substances you will be using in this experiment and complete the following table. Some GHS precaution phrases apply only to laboratory staff handling bulk chemicals or in circumstances where wearing protective clothing and eyewear is not routine. These phrases are displayed in a smaller font and do not need to be copied below.

The web page where you get your safety information has a link to a VIDEO about one of the techniques you will use in this experiment. **You should watch this video before you come to the lab.**

SUBSTANCE	GHS SIGNAL WORD	HAZARDS AND PRECAUTIONS
1 M $\text{Pb}(\text{NO}_3)_2$		
1 M ZnSO_4		
1 M SnCl_2		
1 M CuSO_4		
Pb		
Zn		
Sn		
Cu		

2 M NH ₃		
Fe(NH ₄) ₂ (SO ₄) ₂ ·6H ₂ O(s)		
saturated sodium sulfate		

APPARATUS	RISK	PRECAUTIONS
All glassware including beakers, flasks, funnels, test-tubes.	Glass breakage, cuts from chipped glassware.	Wear safety glasses, covered footwear, lab coat. Inspect all glassware for chips or cracks and take damaged items to the service room for replacement. Do not use damaged glassware.

Pre-lab calculation (compulsory, must be completed to be allowed to start the experiment)

Calculate the mass of Fe(NH₄)₂(SO₄)₂·6H₂O(s) required to prepare 25 mL of 1.0 M Fe(NH₄)₂(SO₄)₂(aq).

Your demonstrator must see your completed pre-lab work and **sign below BEFORE you commence the experiment.**

Demonstrator's signature and date	
-----------------------------------	--

Materials Needed

Cu ²⁺ /Cu reference electrode	1 M Pb(NO ₃) ₂	Pb	150 mL beaker
multimeter	1 M ZnSO ₄	Zn	steam bath
10 mL measuring cylinder	1 M SnCl ₂	Sn	5 cm strip of filter paper
small sample tubes	1 M CuSO ₄	Cu	saturated sodium sulfate
a two holed cap	Fe(NH ₄) ₂ (SO ₄) ₂ ·6H ₂ O(s)	2 M NH ₃	

Technique

Read the instructions below and watch the video online (link on pre-lab safety page) BEFORE you come to the laboratory so you will be prepared to use these instruments. There are details about using a multimeter in the technique section on page 43. You should look over that material before coming to the lab and refer to it in the lab if you are not sure about how the terminals on the multimeter are labeled.

Method

Part A – Galvanic Cells



Perform this experiment as a group. If your bench side has 3 students you will work together. If your bench side has more than 3 students your demonstrator will allocate you into groups. Before you start the procedure, read through what you will be doing and assign tasks to each person so that you can work in parallel where possible (e.g. steps 2, 4, and 5 can be carried out independently in parallel).

- Obtain two tall sample tubes and a white plastic sample tube rack (from the issue room).

2. Prepare 25 mL of approximately 1 M ferrous ammonium sulfate by dissolving the appropriate mass (from pre-lab) of $\text{Fe}(\text{NH}_4)_2(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$ in 25 mL of water (measuring cylinder is adequate) in a beaker. Either use warm water from a hot water tap if available, or warm the solution on a steam bath, to assist the solid dissolving.
3. Cool the solution from 2, then transfer it to one of the sample tubes.
4. Measure 25 mL of 1 M copper (II) sulfate solution into the other sample tube.
5. Clean a long, narrow strip of copper metal and a long iron nail using steel wool on the wooden boards provided.
6. Place the long narrow strip of copper metal in the copper sulfate solution and the iron nail in the ferrous ammonium sulfate solution.
7. Connect the iron electrode to the COM terminal on the voltmeter and the copper to the V Ω A terminal. Make sure the clips grip the metals firmly.
8. Place a strip of filter paper so that each end dips into one of the metal solutions.
9. Obtain a small amount (no more than 2 mL) of saturated sodium sulfate solution in a clean beaker. With a disposable squash pipette, drop saturated sodium sulfate solution onto the filter paper until it is completely wet, but do not add more solution than is necessary.
10. Record the potential difference shown on the multimeter and deduce which electrode is positive. See the 'Technique' section above for information about how the multimeter displays voltages.
11. Remove the filter paper strip and the connections to the multimeter.
12. Place the iron nail in the copper sulfate solution and the copper strip in the ferrous ammonium sulfate solution.
13. Record any visible changes to the appearance of the metals.
14. **Clean the iron nail and copper strip** with steel wool before returning them to where you got them. Do NOT throw the copper and the nail away. Dispose of the solutions in the metal solution waste container in the fume cupboard.

Part B – Electrode Potentials



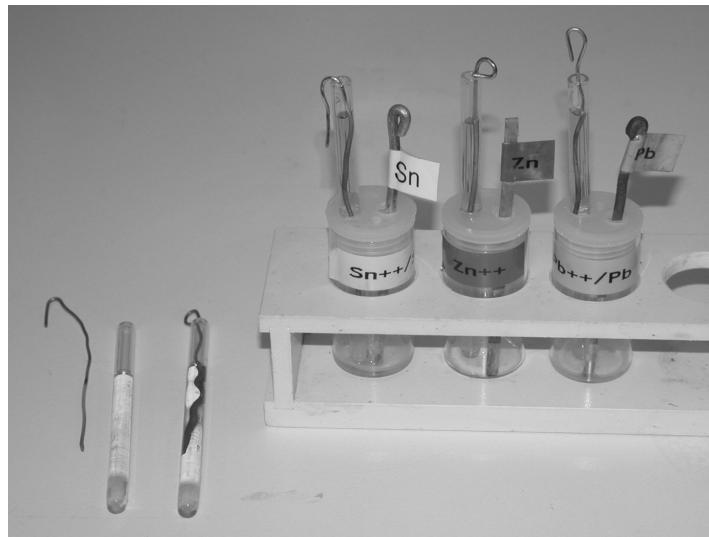
The half cells needed for this experiment have to be prepared by you. At each bench workspaces 1, 2 & 3 will prepare a set of three cells and workspaces 4, 5 & 6 and workspaces 7, 8 & 9 will also prepare a set.

Check that the metals are clean and shiny and free from corrosion. If they are not, they should be cleaned with steel wool in the fume cupboard. Do not scratch the bench surface with steel wool.

1. From the table below, determine the cell that you should prepare, based on your workspace.

Workspace	Half Cell
1 & 4 & 7	$\text{Pb}^{2+} \mid \text{Pb}$
2 & 5 & 8	$\text{Zn}^{2+} \mid \text{Zn}$
3 & 6 & 9	$\text{Sn}^{2+} \mid \text{Sn}$

2. Obtain a small sample tube and a piece of metal appropriate to your workspace. Half fill the tube with the 1 M solution appropriate to your workspace.
3. Insert a two holed plastic cap on the tube, with a copper reference electrode going through the open hole. Insert the piece of clean metal appropriate to your workspace in the second hole. In the picture below the copper reference electrode is shown on the left lying on the bench, and also in place in each of the full cells. As a group, your bench should have a set of cells as pictured below.



4. Attach the lead from the V Ω A terminal of the multimeter to the Cu electrode of the cell and the other lead (from the COM terminal on the voltmeter) to the other metal in the cell. Record in pen the voltage produced by the cell onto your results page. Determine whether the Cu reference electrode is the positive or negative electrode. The multimeter shows a positive reading if the electrode connected to the V Ω A terminal is positive with respect to the electrode connected to the COM terminal. Since you have connected the copper electrode to the V Ω A terminal, a positive reading tells you that the copper electrode is the more positive of the two electrodes in the cell. Disconnect the multimeter from the cell when you have completed your reading.

Make sure you determine the cell voltage and **which half-cell is the more positive** before you disassemble the cell.

Part C – Determination of a Stability Constant [CHEM1031 and CHEM1051 only]

(This part to be done individually)

1. Using the 1.0 M CuSO₄ solution provided, prepare 8.0 mL of 0.20 M CuSO₄.
2. Using the 2.0 M NH₃ solution at your workspace, prepare 8.0 mL of 1.0 M NH₃.
3. Mix the two solution together and half fill a small sample tube. Insert a two holed cap on the tube, together with a copper reference electrode. Insert a piece of clean copper wire in the second hole. Using a multimeter, measure the voltage of the cell.
4. Discard all solutions containing ammonia into the sinks in the fume cupboard.
5. Wash and clean all glassware and apparatus that you have used with detergent provided at the sinks and return them to the common locker or the service room and ask your demonstrator to check your locker.
6. Complete the report and submit to your demonstrator, if possible before leaving the laboratory.

Report Collection

This is the final report you will submit for this course and also for most students this will be your final lab class. Unless advised otherwise, your marked report will be available from the box marked with your demonstrator's name in the School of Chemistry administration area on level 1 of the Dalton Building no earlier than one week after the submission deadline.

School of Chemistry, UNSW

Coversheet for Submission of Individual Report

Report: Galvanic Cells

Circle your course: CHEM1011 CHEM1031 CHEM1051

Student Name (full name): _____

Student Number (e.g., z1234567): _____ Lab. class: day/time: _____

Demonstrator: _____ Workspace: _____

Date and Time Submitted: _____

In preparing this assessment task I have followed the Student Code of Conduct. I certify that I have read and understand the University requirements in respect of student academic misconduct outlined in the Student Code of Conduct and Annexure 1 of the Student Misconduct Procedures. I declare that this assessment item is my own work, except where acknowledged, and has not been submitted for academic credit previously in whole or in part. I acknowledge that the assessor of this item may, for assessment purposes: (1) provide a copy to another staff member of the University; (2) communicate a copy of this assessment item to a plagiarism checking service (such as Turnitin) which may then retain a copy of the assessment item on its database for the purpose of future plagiarism checking. I have retained a copy of this, my assignment, which I can provide if necessary. By signing this declaration I am agreeing to the statements and conditions above.

Student signature and date: _____

Full UNSW policy concerning student misconduct is available at: <<https://student.unsw.edu.au/conduct>>

Feedback (Demonstrator to complete - Skills assessment – tick ALL that apply)

Non-core Skills

Applying chemical principles: correct logic and explanation used appropriate principle(s)

[CHEM1031/51 only] Chemical calculations: correct arithmetic correct method correct sig. fig. on results

Chemical equations: balanced - atoms & charge correct formulae correct states of matter no spectator ions

Describing chemical changes: correct use of terms e.g., precipitate, supernatant

Mastery (Applying chemical principles): clear & concise explanation(s)

Professionalism: answers and discussion in pen locker and contents left clean

Recording observations: correct sig. fig. on measurement(s) in pen (no whiteout) in lab manual

Safety awareness: safety eyewear always on

Time management: worked on report in the laboratory

Comments: _____

Demonstrator signature and date: _____

RESULTS (Results must always be written in pen.)

Part A - Galvanic Cells

Which electrode is the positive electrode? (iron or copper?) _____

Which electrode is the negative electrode? _____

Which electrode is the anode? _____

Which electrode is the cathode? _____

Voltage produced by the iron/copper cell = _____ V

Write the half equations for the reactions at the anode and cathode, and the overall reaction

Anode:

Cathode:

Overall:

Write the specification for this cell using standard notation for electrochemical cells.

Observations of metals in metal ion solutions:

Cu in $\text{Fe}^{2+}(\text{aq})$

Fe in $\text{Cu}^{2+}(\text{aq})$

Which of the two metal in metal ion solutions demonstrates a spontaneous reaction? _____

Write the equation for the overall reaction for this metal in metal ion solution.

When this reaction occurs in a galvanic cell it produces electrical energy. If the metal and metal ion solution were mixed together in one beaker how would this energy be released? (i.e. what form of energy would be produced?).

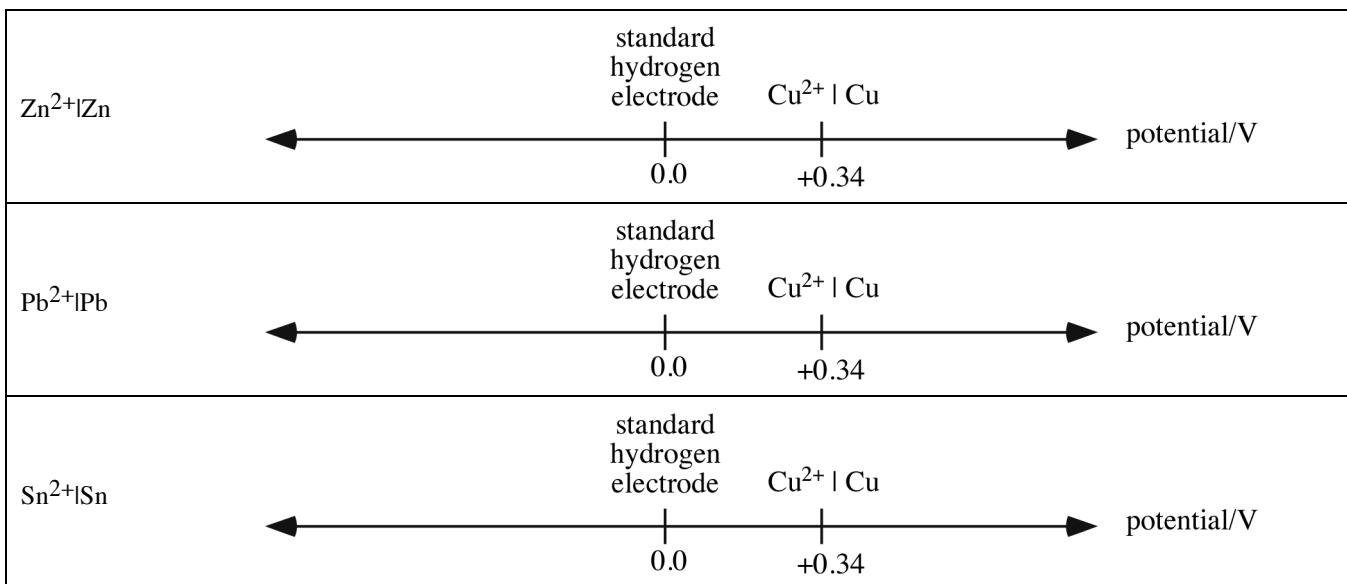
Part B – Electrode Potentials

Measured cell potentials and polarity

Cell constructed from $\text{Cu}^{2+} \text{Cu}$ with:	$\text{Zn}^{2+} \text{Zn}$	$\text{Pb}^{2+} \text{Pb}$	$\text{Sn}^{2+} \text{Sn}$
Measured E_{cell}/V			
Is the copper reference electrode <i>positive</i> or <i>negative</i> relative to the other metal?			

Calculations

For each of the half cells you constructed, indicate the potential of the other half cell with respect to the standard copper electrode on the diagram below. If the copper reference electrode is positive relative to the other metal electrode, the other metal electrode will be to the left of the copper electrode on the diagram, and *vice versa*. (E° for $\text{Cu}^{2+}|\text{Cu} = +0.34 \text{ V}$)

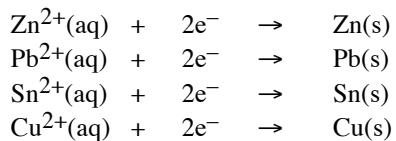


Use the diagrams above to deduce the potential of each of the metal ion/metal half cells with respect to the standard hydrogen electrode. These are the standard reduction potentials for the metals.

Show your working for calculating E° for the $\text{Zn}^{2+}|\text{Zn}$ half cell.

	$\text{Zn}^{2+} \text{Zn}$	$\text{Pb}^{2+} \text{Pb}$	$\text{Sn}^{2+} \text{Sn}$
Experimental E°/V (referred to standard $\text{H}^{+} \text{H}_2$ half cell)			
Literature E°/V (From SI Chemical Data)			

From your results, arrange these reductions in order of decreasing reduction potential:



Highest (most positive) reduction potential	
Lowest (least positive) reduction potential	

From ALL the species listed above (metals and metal ions), select:

- a) the best oxidant b) the best reductant

Provide a brief justification for your choice of best oxidant and best reductant.

The Nernst equation in Analytical Chemistry [CHEM1031 only]

Briefly describe an application of the Nernst equation in the measurement of the concentration of a species in solution, giving a specific example.

The Nernst equation in Biochemistry [CHEM1051 only]

There is a potential difference between the inside and outside of a resting nerve cell because of differing concentrations of ions in the interior and exterior of the cell. If the concentration of potassium ions is 10 mM outside the cell, and 100 mM inside the cell, what is the potential difference across the cell membrane at 37 °C according to the Nernst equation?

Part C – Determination of Stability Constant [CHEM1031 and CHEM1051 only]

Cell	Voltage
Cu $[\text{Cu}(\text{NH}_3)_4]^{2+}$ (aq), NH_3 (aq) Cu^{2+} (aq) Cu	

Using the cell voltage calculate the value of the stability constant, K_{stab} . Hints:

1. Calculate the concentrations of $[\text{Cu}(\text{NH}_3)_4]^{2+}$ (aq) and NH_3 (aq) in the expression for Q from the initial concentrations of Cu^{2+} and NH_3 which you mixed in step 3, (after allowing for mutual dilution) assuming this reaction goes to completion: $\text{Cu}^{2+}(\text{aq}) + 4\text{NH}_3(\text{aq}) \rightarrow [\text{Cu}(\text{NH}_3)_4]^{2+}(\text{aq})$.
2. The concentration of Cu^{2+} in the expression for Q is the concentration in the reference electrode.

1. Calculate the concentrations of NH_3 and $[\text{Cu}(\text{NH}_3)_4]^{2+}$ in the left-hand half cell. These are the concentrations referred to as x and y respectively in the introduction. Do not forget to allow for the mutual dilution which happens when you mix solutions.

$$[\text{Cu}(\text{NH}_3)_4]^{2+} / \text{mol L}^{-1} = \dots$$

$$[\text{NH}_3] / \text{mol L}^{-1} = \dots$$

2. Use the concentrations from step 1 and the known concentration of Cu^{2+} (aq) in the reference electrode to calculate Q .

$$Q = \dots$$

(Calculation continues on the next page.)

3. Use the (re-arranged) form of the Nernst equation given in the introduction to calculate $\ln K_{\text{stab}}$ and hence K_{stab} from the measured cell potential and the value of Q calculated above.

$K_{\text{stab}} = \dots$

WHAT'S NEXT?

We hope you enjoyed the laboratory component of your Chemistry A course.

We welcome your feedback and hope you will continue to study chemistry in future semesters. In semester 2 the School of Chemistry offers CHEM1021 – Chemistry B: Elements, Compounds and Life and, in semester 2 only, CHEM1041 – Higher Chemistry B and CHEM1061 – Higher Chemistry (Medicinal) B, all of which build upon the material in Chemistry A and expand your knowledge of chemistry into the areas of:

- Structure determination – how do chemists know what the shape of a molecule is? How do you determine the distances between atoms in a molecule or between ions in a crystal? How can you tell which atoms are bonded together in a molecule? How can you determine the composition of mixtures (even on the surface of another planet)?
- Rates of reactions – what determines how fast chemical reactions occur? Why are some reactions complete in nanoseconds while others take days or years? How does temperature affect the rate of a reaction? (And why is this important in living organisms, and in cooking and food preservation?) How long are drugs effective in the human body?
- Inorganic chemistry – the properties and reactions of non-organic compounds. Why are huge quantities of sulfuric acid and ammonia produced each year? Why are metallic elements so important to living organisms? (Did you know you have about 0.1 g of copper in your body?) What challenges are there in creating more efficient batteries and in storage of energy?
- Organic chemistry – why is the chemistry of carbon so special? Find out how chemists craft new molecules for applications as diverse as pharmaceuticals, protecting plants from insects, stopping barnacles attaching to ship hulls, and making better contact lenses.

If you have any questions about these courses, or just want more information, please contact our Student Support Manager (Steve Yannoulatos, Dalton 104) or Anne Ayres (Dalton 105) or email firstyearchem@unsw.edu.au.

Cool Research in the School of Chemistry

A/Prof. Stephen Colbran and his group are studying
Chemical and Biological Catalysis:
Turning light into electricity

Dye-sensitised solar cells (DSSCs) offer the possibility of converting everyday building surfaces — roofs, walls, even windows — into sun-driven electricity generators. The problem is that at present they are all too costly or too low efficiency. Working with Dr Andrew McDonagh at UTS, the group characterised the electrochemistry and spectroscopy of new ruthenium phthalocyanin-based dyes. Studies of the new dyes in DSSCs revealed important interplays between dye redox-properties, sizes and coverage and light-driven electricity generation efficiencies.

So how does Nature employ metal centres to catalyse complex multi-electron processes? A/Prof. Stephen Colbran and his group focus on biomimetic catalysts, one of them being depicted below.

For more information about this research see:
<http://www.chemistry.unsw.edu.au/our-school/staff-profiles/stephen-colbran>

