CHEMISTRY FOR FUN
Demonstrating chemistry experiments

Author
Tomaž Ogrin, Jožef Stefan Institute, Slovenia; tomaz.ogrin@ijs.si
Languages available
English, Slovenian

Summary
The activity aims to stimulate students to demonstrate experiments to other students together with the explanation of some related chemistry. Experiments will also offer an opportunity for research work in changing parameters and reviewing procedures.

Activity type D
Use of scientific knowledge and understanding to solve problems
Working in teams to solve problems
Resource/budget management

Techniques
Laboratory techniques

Field
Chemical experiments
Education

Time:
Practical lessons: 100 minutes
Theory lessons: 80 minutes
Out of class time: 90 minutes

StandardBase procedures
Determination of pH and conductivity using electrochemical techniques

StandardBase techniques
Potentiometry
Conductimetry

Other resources
http://chemistry.about.com/od/homeexperiments/Chemistry_Experiments_You_Can_Do_at_Home.htm (An experiment ideas site together with methods.)
http://chemistry.about.com/od/chemistrydemonstrations/ss/bluebottle.htm (Blue bottle experiment, also recommended, methods are straightforward.)
http://chemistry.about.com/od/demonstrationsexperiments/ss/smokebomb.htm (Smoke bomb - out of class activity.)

Student's document

CHEMISTRY FOR FUN: Demonstrating chemistry experiments

Vocational chemistry students can go on to obtain jobs in schools – and different types of school, from primary to university. You work as an assistant to a chemistry teacher. Your job is to prepare experiments for chemistry teachers to use in the chemistry courses they teach.

Your brief
You will work in a group.
1. Your group is asked to prepare and perform an experimental part of a lecture, instead of a teacher, together with short explanations of the experiments, to a class of pupils or students. There are four lectures (courses, issues) planned. Each presentation should be about 30 minutes. Experiments and presentations should be interesting and attractive to the students.

2. In addition, your group is invited to perform experiments at other schools. This time, you have to have a budget to cover expenses. You are asked to calculate the budget, including transport expenses to and from other locations.

3. Chemistry course contents include the following areas of general chemistry: acids and bases, buffers, chemical redox reactions and rate of chemical reactions.

4. A presentation at a course is composed of three parts; these are to be shared among your group (use **Student sheets)**:
a) A brief introduction to the issue
b) A demonstration of experiments
c) Explanations of the phenomena seen

5. The task is open-ended, which means you can add your own ideas to the activity.

Your investigation
Overview of tasks
1. Prepare an introduction, the experimental part and the explanation for each of the courses, which have the following titles:
a. Acids and bases
b. Buffers
c. Redox reactions
d. Chemical rate
Use all four **Student sheets**. Adapt to suit the knowledge-level of the audience and use different text, written by yourself.

2. Make a list of chemicals and accessories for each experiment.

3. Estimate expenses for chemicals, accessories, equipment, transport by car (taking into account the use of a private car, school car or rent-a-car) and the labour cost. Transport is to be taken as price per kilometre from starting point to location and back.

4. Calculate the required budget for the demonstrations for all four courses.

5. Estimate how distance in kilometres influences the proportion of the budget needed for transport.

6. You will need the CAS numbers (CAS = Chemical Abstract Service) and MSDS numbers (MSDS = Material Data Safety Sheets). For additional chemicals, visit: http://msds.chem.ox.ac.uk/#MSDS

7. Measure the time needed to conduct an experiment, for all the experiments.

8. A risk assessment must be carried out before starting practical work.
Risk assessment is made using MSDS data. It is suggested that you use data from http://msds.chem.ox.ac.uk/#MSDS

Task 1
Using **Student sheet***: Acids and bases*
- Give a short introduction to acids and bases.
- Perform the experiment showing change of colour on adding acid or base to pure water. Use pH indicator.
- Add the opposite, acid or base, to an acidic or basic solution until the colour is changed back.

Task 2
Using **Student sheet***: Chemical buffers*
- Give a short introduction to acid-base buffer solutions.
- Perform the experiment showing the delayed change of colour after the addition of acid or base, when buffers are present in the solution.
- Add the opposite, acid or base, to an acidic or basic solution until the colour is changed back.
- Explain the observations.

Task 3
Using **Student sheet***: Oxidation–reduction reactions*
- Give a short introduction to oxidation – reductions.
- Perform experiments with H_2O_2 - following the idea in the text.
- Explain the observations.

Task 4
Using **Student sheet***: Chemical rate*
- Give a short introduction to chemical change.
- Perform the experiment.
- Explain the observations.

Task 5
Budget and resource managing
- Estimate expenses for chemicals, accessories, equipment, transport by car and a labour cost.
- Calculate the required budget for the performance of all four courses.
- Estimate how distance in kilometres influences the proportion of the budget needed for transport.

Your findings

- Write down some tips, based on your experience with the experiments.
- Discuss your experience with your group.
- Evaluate your own performance first and then follow this with a class evaluation.
- Prepare a poster as a group, including some photos.
- Write a short report on the tasks you completed.

Student self assessment

This activity was about using scientific knowledge and understanding to solve problems, working in teams to solve problems and resource and budget management. For each case, write about 50-100 words explaining:

- What scientific skills and knowledge you used and how effectively you used them.
- How you worked as a team, giving examples of how you supported others in your group or were supported by them.
- How useful you think it was to work out the cost of demonstrations.

You may decide to use alternative forms of self assessment.
The use of the file *Evaluating skills.doc* from the **ProBase** website is suggested.

Student sheet: Acids and bases

Below is an example of a brief introduction for your lecture about acids and bases (it could be much shortened). Using this text and links, prepare your own introduction.

Scope
Acids and bases are very important classes of compounds, especially in aqueous solutions. They are also a part of our bodies (for example: HCl in the stomach). Here, we are talking about acidic, neutral and basic aqueous solutions. In these experiments, the role of pH indicators - specific compounds chosen to test the acidity of solutions - will be demonstrated.

Principle
The well known measure of acidity is pH. Acidic solutions have pH lower than 7, basic solutions higher than 7. Solutions with pH = 7 are neutral. pH denotes a mathematical expression for the H_3O^+ (hydronium) or simplified H^+ (proton, hydrogen nucleus) ions concentration. Greater H^+ concentration means more acidic solution, but lower pH. The opposite holds for basic solutions.

How are acids and bases connected to pH and to hydrogen ion concentration? For our purpose, we will use the first, very useful definition of acids and bases, introduced by Arrhenius: Acids are compounds that release H^+ ions into a solution. Bases are compounds that release OH^- ions into a solution.

But it all begins with water. Water is mainly composed of H_2O molecules. Some of them, a very small quantity, are decomposed to H^+ and OH^- ions. Hydrogen ions H^+ (hydrogen nucleuses, protons) are never alone in a solution, because they are very small, charged species. They are tied to water molecules, forming H_3O^+ (hydronium) ions and ions with more H_2O molecules attached (minimum three more). In a very simplified picture, we could use H^+ notion, instead of H_3O^+. The concentration of hydronium and hydroxyl ions are equal in pure water, making it neutral, not acidic nor basic.

When we add some acid to the water, for example acetic acid (as in vinegar), a solution becomes acidic; acetic acid releases some H^+. It means that there is more H^+ (H_3O^+) than OH^- after the addition of vinegar. The opposite is true when we add some basic material, for example washing powder.

So, how can we tell if, for example, a solution is acidic? We know when we eat a salad, but tasting everything is not wise in chemistry.

Some compounds exist, known as pH indicators, which tell us about pH. They are sensitive to concentrations of H^+ ions. They change colour when the concentration of H^+ changes. They have a definite range of H^+ concentrations in which they change colour. One well known indicator is in red cabbage. It is what gives a cabbage its reddish-violet colour. You can actually make your own indicator from it.

Chemistry handbooks and textbooks can be used to find the range of change for many indicators. Also see the Internet sites given below.

http://en.wikipedia.org/wiki/Methyl_orange
http://en.wikipedia.org/wiki/Phenolphthalein
Try other indicators:
http://antoine.frostburg.edu/chem/senese/101/acidbase/indicators.shtml

Procedure

- Into two test tubes, pour distilled water to half of their volume.
- Add pH indicator (in solution), phenolphthalein to one tube and methyl orange (sodium salt) to the other, and shake.
- Add three to four drops of HCl into the test tube with methyl orange and shake.
- Add three to four drops of NaOH into the test tube with phenolphthalein and shake.
- Now, add six to seven drops of NaOH into the test tube with HCl and shake.
- Add six to seven drops of HCl into the test tube with NaOH and shake.
- Comment to the audience on the changes of colour.

Hint

Prepare the experiment before your lecture by investigating several indicators, acids and bases and make a choice for the demonstration. Also select the glassware you will need (test tubes, conical flasks, etc.). Determine quantities (drops and concentrations) of each acid or base to be added to a solution to attain a colour change, several times.

Safety aspects

While ionic reactions do not usually produce much energy, some cases need to be stressed. On dilution of concentrated acids (for example 98% sulfuric acid H_2SO_4) and bases (for example pellets of NaOH), much heat is produced (heat of dilution or dissolution). As a safety rule, always pour a small quantity of concentrated acids into a large volume of water and not vice versa. Wear safety goggles and gloves, together with an apron. Make dilutions in a fume hood.

Student sheet: Chemical buffers

Below is an example of a brief introduction for your lecture about chemical buffers (it could be much shortened). Using this text and links, prepare your own introduction.

Scope
An experiment is performed in which we:
- show changes of the acidity of solutions when buffers are present in a solution and
- compare these changes to changes when only a strong acid or base is present without buffers in a solution.

Principle
Buffers absorb the effect of an acid or base added to a solution so that the acidity is not changed significantly, unless we add a huge amount of acid to it. This is valid to some extent, depending on the capacity of the buffer. Without a buffer, on adding an acid or base to the solution, its acidity would change to a much greater extent. Buffers are mostly mixtures of two solutions. An acidic buffer solution contains a weak acid plus one of its salts with a strong base (more precisely a conjugate base in a Brönsted definition). On the basic side, the buffers used are composed from a weak base and one of its salts with a strong acid.

A brief explanation of the term buffer can be given first. It could also be given after an experiment, when the audience will have seen that there is not the same change of colour when about the same quantity of acid or base was added.

It could be mentioned that our body keeps the acidity of blood very constant using a series of buffers. If our blood acidity changes, our acid-base balance is disrupted, leading to acidosis (our blood is too acidic) or alkalosis (our blood is too alkaline).

Procedure
- Hint: use the same system that you devised for **Student sheet**: *Acids and bases.*
- Into two test tubes, pour distilled water to half of their volume.
- Add 4-5 mL of buffer.
- Add pH indicator (in solution), phenolphthalein to one tube and methyl orange (sodium salt) to the other, and shake.
- Add three to four drops of HCl into the test tube with the methyl orange and shake.
- Then add additional drops of HCl until the colour changes.
- Add three to four drops of NaOH into the test tube with the phenolphthalein and shake.
- Then add additional drops of NaOH until the colour changes.
- Comment to the audience on the changes of colours and the difference in the quantity of the acid or base to be added to change the colour when buffers are present.

Hints
Prepare for your lecture by investigating the buffer system and the quantities of acid and base that needs to be delivered to a buffered solution to show the effect as clearly as possible.

Suggestion: search for ‘acidosis’; look at, for example:
http://www.wrongdiagnosis.com/a/acidosis/basics.htm#howserious
Or type in ‘blood pH’:
http://www.chemistry.wustl.edu/~edudev/LabTutorials/Buffer/Buffer.html
http://www.bbc.co.uk/dna/h2g2/A8819652

Look for other buffered systems.
To find out about buffers, look at:
D.A.Skoog, D.M.West, F.J.Holler: *Fundamentals of Analytical Chemistry*, 7th Ed., Saunders College Publ., 1996, p.200
J.N.Butler, *Solubility and pH calculations*, Adison-Wesley Publ.Co., 1964
E.J.Margolis, *Chemical Principles in Calculations of Ionic Equilibria*, The MacMillan Co., N.Y., 1966

Student sheet: Oxidation–reduction reactions

Below is an example of a brief introduction for your lecture about oxidation – reduction reactions (it could be much shortened). Using this text and links, prepare your introduction yourself.

Scope
Oxidation-reduction, or redox reactions, are characterised by electron transfer (or their reorganisation, redistribution), while in acid–base reactions, it is the protons that move. There is some analogy between both types of reaction. Both involve two conjugate pairs (Brönsted): two acids, two bases in an acid–base type, and two oxidised forms and two reduced forms in an oxidation–reduction type.

About terms: the *oxidising agent* gains or, rather pulls, electrons; the *reducing agent* loses them. We introduce oxidation numbers as a tool to follow electron transfer. The periodic system is used to assign oxidation numbers to atoms.

Principle
Oxidising agents differ in their strength of oxidising ability (pulling electrons) and reducing agents differ in their ease of losing electrons (think about ionisation energies). We can rank them in a list, according to their strength. The elements and their ranking can be identified from their position in the Periodic Table: the strongest oxidising agent among the elements is fluorine and the other halogens are also strong; the strongest reducing agent is lithium and the other metals in the first column of the Periodic Table are also strong (alkalis).

Electron transfer among atoms and molecules can create a current of electrons through wires and appliances (light bulbs for instance). Technology utilises these redox reactions in batteries and fuel cells. The voltage achieved depends on the strength of the oxidation and reduction ability of the constituents. Thus, the voltage of modern lithium batteries is 3 V, while others are only 1.5 V. There are published tables that give the electrode potential for various combinations of oxidants and reductants.

Safety aspects
Redox reactions usually produce much heat in a short time, which could lead to explosions or ignition. The well known reaction of sodium metal with water produces hydrogen which self-ignites and burns with small explosions.

Safety goggles and gloves are essential together with a laboratory apron/coat.

Handling concentrated (20%–30% and over) hydrogen peroxide needs special care. Look at the MSDS report for safety aspects at:
http://www.bu.edu/es/labsafety/ESMSDSs/MSHydPeroxide.html

Hydrogen peroxide slowly releases oxygen. Bottles (usually not made of glass) with concentrated H_2O_2 have tiny openings in the stopper to release oxygen and so prevent pressure build-up in the bottle.
Decomposition of hydrogen peroxide speeds up considerably if certain catalysts (metals, metal oxides) are present – building up a high pressure when confined in a bottle, leading to possible explosion.

Experiments with H_2O_2

Hydrogen peroxide is a molecule composed of two oxygen atoms, which are bonded to each other, and two hydrogen atoms, bonded to each oxygen atom.

When hydrogen peroxide simultaneously undergoes both oxidation and reduction, we call such a reaction disproportionation. It is easily disproportionated (decomposed) by catalysts (such as metal oxides or saliva-containing enzymes which are bio-organic catalysts). Products of decomposition are water and oxygen.

When a reducing agent is present, hydrogen peroxide acts as an oxidising agent. When stronger oxidising agents are present, hydrogen peroxide acts as a reducing agent.

1. Experiment: H_2O_2 disproportionation

Hydrogen peroxide decomposes into water and oxygen with the evolution of considerable heat:

$$2\ H_2O_2 \rightarrow 2\ H_2O + O_2 \qquad \Delta_r H = -\ 196.2\ \text{kJ/mol}$$

At room temperature, the decay rate is so immeasurably slow that hydrogen peroxide is effectively stable (metastable) in both pure form and in solution. As mentioned, the reaction rate is greatly increased by catalysts (e.g. heterogeneous catalysts such as finely-dispersed silver, gold, platinum, manganese dioxide or dust particles, or homogeneous catalysts such as I^-, IO_3^-, OH^-, Fe^{3+} or Cu^{2+}). In their presence, a violent evolution of oxygen occurs at room temperature and, in highly concentrated solutions, even explosive decay because of excessive heat development.

Procedure

Into a 2 L beaker, pour 200 mL 30% H_2O_2. Add 5–10 mL liquid soap (contains surface active ingredient – tenside) and mix both liquids. Tensides are compounds that lower surface tension. Then add 20 mL of 1 mol/dm^3 KI. Hydrogen peroxide decomposes and a large quantity of foam is formed. It is wise to have a pad or polyethene (PE) foil under the apparatus.

Developed oxygen in the foam is detected by using a wooden stick heated until glowing red. The stick will glow brightly in the foam, giving the impression of glowing foam. The reason is oxygen-rich foam (look at the equation).

Investigate the proportion of KI and peroxide to find the optimum amount of foam for your set-up.
Start the search with a fixed amount of 30% H_2O_2 (for example 200 mL) and a fixed amount of KI (for example 20 mL of 1 mol/dm^3). Vary the amount of liquid soap 5 mL, 10 mL, 15 mL, 20 mL. Then, keeping the amount of liquid soap constant at the figure that gave the most foam, vary the amount of KI at 15 mL, 20 mL and 25 mL. After that, keep the amount of liquid soap and KI constant at the figure that gave the most foam and vary the amount of hydrogen peroxide at 150 mL and 250 mL. Write down the results in a table.

2. Experiments: H_2O_2 as oxidant

2.1. Experiment: From colourless solution to a strong red one: oxidation of Fe^{2+} to Fe^{3+} with hydrogen peroxide, and successive reaction to red SCN^- complex. Investigate how to present an experiment and make a draft of the presentation.

Procedure
A recipe for the experiment is as follows:
1. 2-3 mL of 0.1 mol/dm^3 $Fe(NH_4)_2(SO_4)_2$ solution is poured into a test tube.
2. Add some drops of 2 mol/dm^3 sulfuric acid.
3. Add a point-of-a-knife amount of Fe powder.
4. Decant the solution into another test tube.
5. To this solution add 1–2 drops w = 1% NH_4SCN solution.
6. Then add some drops of w = 3% H_2O_2 solution.

The solution will become intensively red. Hydrogen peroxide oxidises Fe^{2+} to Fe^{3+} giving a red colour at a reaction with SCN^- ions. H_2O_2 is reduced to water.

2.2. Experiment: From black to white: black PbS turns white $PbSO_4$ by oxidation with hydrogen peroxide. Investigate how to present an experiment and make a draft of the presentation.

Procedure
A recipe for the experiment is as follows:
1. Pour into a test tube about 2 mL 0.1 mol/dm^3 $Pb(NO_3)_2$.
2. Add about 5 mL 0.1 mol/dm^3 Na_2S to get a black precipitate of PbS.
3. Decant the solution over the precipitate.
4. Wash out the precipitate with hot distilled water several times.
5. To the precipitate add some mL of w = 3% H_2O_2.

The black colour disappeared: PbS was transformed to white precipitate $PbSO_4$ by the action of hydrogen peroxide, which was reduced to water.

3. Experiments: H_2O_2 as a reductant
3.1. Experiment: Decolouration of a violet solution: potassium permanganate reduction with hydrogen peroxide. Investigate how to present an experiment and make a draft of the presentation.

Procedure
A recipe for the experiment is as follows:
1. Pour into a test tube 3 mL 0.02 mol/dm^3 $KMnO_4$ solution.
2. Add some drops 2 mol/dm^3 H_2SO_4 solution.
3. Add drops w = 3% H_2O_2 until discolouration.

Violet MnO_4^- ion was reduced to nearly colourless Mn^{2+}. H_2O_2 oxidised to oxygen, bubbling through the solution.

3.2. Experiment: Silver from a solution: silver oxide reduction with hydrogen peroxide. Investigate how to present an experiment and make a draft of the presentation.

Procedure
A recipe for the experiment is as follows:
1. Pour into a test tube 2 mL 0.1 mol/dm^3 $AgNO_3$ solution.
2. Add some drops of 1 mol/dm^3 NaOH solution.
3. Carefully decant the solution over the precipitate.
4. Wash out the precipitate with distilled water several times.
5. Decant the solution over the precipitate.

6. Pour into a test tube with the precipitate 2 – 3 mL w = 3% H_2O_2 solution.
7. Prove that oxygen comes out of the test tube with a red glowing wood chip.

The precipitate Ag_2O was reduced with hydrogen peroxide to Ag. H_2O_2 was oxidised to oxygen.

3.3. Experiment: Traffic Light reaction

A conical flask containing a pale yellow solution is gently swirled. The solution turns red. The flask is shaken and the solution turns green. Investigate how to present an experiment and make a draft of the presentation.

Procedure

A recipe for the experiment and parameters variation is as follows:
1. Make a solution A: 3 g dextrose (glucose) and 5 g NaOH dissolve in 250 mL distilled water.
2. Make 1.0% indigo carmine indicator solution.
3. Pour 50 mL of A into a 250 mL conical flask.
4. Add 5–10 mL indicator solution. The result should be a light yellow solution.
5. Stopper the flask.
6. Gently swirl the flask to produce the red colour. If the red colour does not persist, increase the number of drops of indicator.
7. Quickly shake the flask to produce the green colour.
8. On standing, the solution becomes yellow again.

http://www.rsc.org/Education/EiC/issues/2005July/Exhibitionchemistry.asp
L.R.Summerlin, J.L.Ealy, *Chemical demonstrations, Vol.1*, 2nd Ed., American Chemical Soc.,1988, p. 111
http://www.youtube.com/watch?v=QkUKlxzNbow

Student sheet: Chemical rate

Below is an example of a brief introduction for your lecture about chemical rate (it could be much shortened). Using this text and links, prepare your introduction yourself.

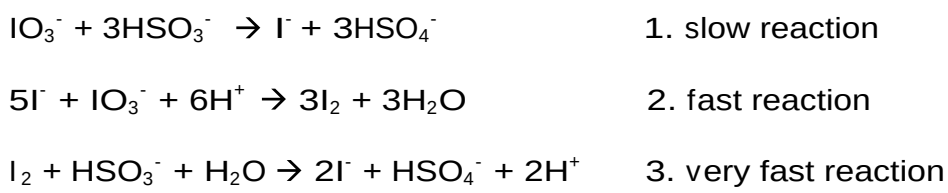
Scope
These types of reactions (clock reactions) give visual results after a definite time. This time could be managed by changing the reaction and ingredient parameters.

Iodine Clock reaction
You could tell the audience in advance the time when the solution will suddenly turn dark blue. They will follow your forecast on their watches to the second. Of course, you must test the time when you prepare the experiment, bearing in mind the influence of the concentrations of reactants (including how accurately the solutions were made and the volumes delivered) and the temperature of the solutions.

This is what chemical reaction rates depend on. The influence of catalysers and/or inhibitors on chemical rate is not involved in this experiment. To avoid any such influence, your laboratory work area and practices should be clean.

It is good practice to prepare stock reactant solutions for more experiments than you think you will carry out.

Explanation
The *Iodine Clock* reaction is based on the following reactions:



$IO_3^- + 3HSO_3^- \rightarrow I^- + 3HSO_4^-$ 1. slow reaction

$5I^- + IO_3^- + 6H^+ \rightarrow 3I_2 + 3H_2O$ 2. fast reaction

$I_2 + HSO_3^- + H_2O \rightarrow 2I^- + HSO_4^- + 2H^+$ 3. very fast reaction

When all HSO_3^- ions are used up, iodine concentration from the second reaction rises and the solution becomes deep blue, because of the reaction between starch (amylose) and iodine.
http://www.elmhurst.edu/~chm/vchembook/548starchiodine.html

Investigate how to present an experiment and make a draft of the presentation.

Procedure
Preparation of solutions:
For solution A:

- 0.02 mol/dm^3 KIO_3: Dissolve 4.3 g KIO_3 in distilled water in a 1 L volumetric flask, fill it up to 1000 mL and homogenise the solution.
- 0.2 g/100 mL soluble starch solution: mix 0.5 g soluble starch with 10 mL cold distilled water, pour the suspension into 240 mL boiling hot distilled water. To preserve, add 2–3 mg HgI_2. If you are going to use it fresh, the preservative is not needed.

For solution B:

- 0.01 mol/dm^3 Na_2SO_3: mix 39 mL 1 mol/dm^3 H_2SO_4 and 10 mL ethanol into 500 mL distilled water, add 1.26 g Na_2SO_3 (or 2.52 g $Na_2SO_3.7H_2O$) and dilute to 1000 mL.

A recipe for the experiment and parameters variation is as follows:

1. Make solution A: Into a 400 mL beaker, measure, using a volumetric pipette, 50 mL 0.02 mol/dm^3 KIO_3 solution and 10 mL starch solution.
2. Make solution B: Into a 400 mL beaker, measure, using a volumetric pipette, 50 mL freshly prepared 0.01 mol/dm^3 Na_2SO_3 solution.
3. To the beaker with A, add 300 mL distilled water and pour all at once into solution B; measure the time for the blue colour to appear (about 2 minutes).
4. As in 3. but with the addition of only 200 mL distilled water. Measure time in seconds.
5. As in 3. but with the addition of only 100 mL distilled water. Measure time in seconds.
6. Record the temperature of the solutions and water before mixing. Keep the temperatures of all solutions constant before mixing.

The results of experiments should be drawn up in a table:

Water added (mL)	**Volume proportion (KIO_3+Na_2SO_3) : H_2O**	**Reaction time (s)**	**Temperature H_2O (^{0}C)**	**Temperature A + B (^{0}C)**
300	1 : 3			
200	1 : 2			
100	1 : 1			

Teacher's document

Overview
This activity promotes a self supported approach as students/group of students, search on the web, develop and then actually demonstrate, with accompanying explanation, some basic chemistry (and physics) to a class of younger students. Only laboratory and other (computer, materials, etc.) facilities are made available to students.

Carrying out most of the reactions involved – and demonstrating them to younger students - does require some basic knowledge and some previous experience in handling chemicals. It is suggested that practical knowledge about handling chemicals is tested too.

Developing students' skills
If, at any stage, you feel that students are not going to meet the objectives of the task, you could give them some specific support and suggestions.

Organising and managing the class
Organising groups
- Groups of two or three students are formed to conduct their own experiments.
- Time to conduct an experiment should be measured when first trying it out. Then experiments can be fairly distributed between students and groups.
- Sub-activities are planned so that the whole activity can be completed in the time allowed, 270 minutes.

Allow at least one computer, with Internet access, per group and laboratory facilities.
Students conduct experiments first in a laboratory, later in a class.

The activity
- The field in chemistry suggested with this activity is an introductory chemical course, in the area of chemical reactions, ionic (acid–base) and redox reactions.
- The sub-activities, where the student becomes lecturer, are structured to include a brief introduction to the topic, followed by a practical demonstration of the experiment.
- The suggested experiments:
  - changes of colour
  - chemical buffers
  - hydrogen peroxide
  - clock reactions
- The teacher should decide on purchase of chemicals, if they are not already available in the laboratory. So teachers should see a list of materials and the planned apparatus set–up, at an early stage. It is also for the teacher to decide whether or not a specific experiment is allowable.
- Laboratory and other (computer, materials, etc.) facilities should be available to students.
- Prices for likely expenses should be listed for students (transport, etc.).

Session plan for teachers
Practical lessons: 100 minutes
Theory lessons: 80 minutes
Out of class time: 90 minutes

- Introducing the topic and checking prior knowledge arranging students into groups (20 minutes)
- Preparation time for students to discuss in a group, plan how to approach the task and allocate work (20 minutes)
- Time for searching the web, prepare and conduct experiments (100 minutes class time plus 90 minutes extra time outside of lessons may also be needed)
- Giving presentations (20 minutes)
- Discussing results, drawing conclusions (20 minutes)

Note: Time for searching the web for experiments could be extended by about 90 minutes out of class time.

Technical notes

Preparation

Students prepare a list of chemicals and equipment for experiments by themselves and agree it with the teacher. It is expected that students know how to prepare 1 mol/L solutions and pH indicator solutions (or these may already be prepared as standard in a general or analytical chemistry laboratory).

Buffers

Buffers are needed for checking the accuracy of pH meters. These may be buffers made from acetic acid and its sodium salt for acidic range and ammonia-ammonium chloride buffer for basic range.

Safety aspects

Be aware of the safety aspects that are essential to consider. This is especially important in the case of the redox reactions. It is essential to conform to safety regulations. This requires vigilance throughout the activity.

Equipment and materials

Materials and their CAS numbers

As well as being a reference for CAS numbers for many other chemicals, other safety data can be found at:
http://msds.chem.ox.ac.uk/HY/hydrochloric_acid.html

Name	Grade	CAS-No.
Indicators		
phenolphthalein		77-09-8
methyl orange (Na salt)		547-58-0
bromothymol blue		76-59-5
indigo carmine		860-22-0
universal (Yamada)		
Other chemicals		
acetic acid	Reagent	64-19-7
α-d-glucose	Reagent	50-99-7
ammonium chloride	Reagent	12125-02-9
ammonium hydroxide	Reagent	1336-21-6
ammonium thiocyanate	Reagent	1762-95-4
ethanol (96% by volume)	Reagent	64-17-5
hydrochloric acid, conc.	Reagent	7647-01-0
hydrochloric acid HCl (1 mol/L)	Reagent	
hydrogen peroxide, 30%	Reagent	7722-84-1
iron (powder)	Reagent	7439-89-6
iron(II)ammonium sulphate hexahydrate	Reagent	7783-85-9
lead(II) nitrate	Reagent	10099-74-8
liquid soap		
potassium iodate	Reagent	7758-05-6
potassium iodide	Reagent	7681-11-0
potassium permanganate	Reagent	7722-64-7
silver nitrate	Reagent	7761-88-8
sodium acetate	Reagent	127-09-3
sodium hydroxide	Reagent	1310-73-2

sodium hydroxide (NaOH 1 mol/L)	Reagent	
sodium sulphite	Reagent	7757-83-7
disodium sulphide	Reagent	1313-82-2
starch	Reagent	9005-25-8
sulfuric acid, conc.	Reagent	7664-93-9
sulfuric acid (1 mol/L)	Reagent	7664-93-9

Equipment

- laboratory balance
- test tubes
- delivery tubes for drops
- graduated cylinders for 5 mL and 50 mL
- volumetric pipette 50 mL
- 100 mL beaker
- 300 mL beaker
- 400 mL beaker
- 2 L beaker

TESTING PRIOR KNOWLEDGE

Select the correct answer.

1. Some substances are called electrolytes. We distinguish between strong and weak electrolytes. Which one on the list is a strong electrolyte in an aqueous solution?

a. HCOOH
b. CH_3COOH
c. H_2S
d. NaCl

2. Acids and bases belong to the class of electrolytes. Electrolytes conduct electric current. For this to happen, species which are called ions must be present in a solution. Which of the listed species is an ion?

a. H_2O
b. NO_3^-
c. CH_3COOH
g. CH_4

3. Which of the listed chemicals is the acid in an aqueous solution?

a. NaOH
b. $Ca(OH)_2$
c. HCl
d. NaCl

4. The pH of an acid solution is:

a. 7
b. higher than 7
c. lower than 7
d. 14

5. A buffer is a mixture (solution) of a weak acid and its salt with a strong base or a weak base and its salt with a strong acid. Buffers are used to keep the pH of a solution constant. Several buffers are in our body to keep the pH of blood constant at 7.4. Find one of the buffers in the list below.

a. acetic acid – sodium acetate solution
b. hydrochloric acid – sodium chloride solution
c. potassium bromide – sodium chloride solution
d. lithium fluoride – sulfuric(VI) acid

7. In oxidation–reduction (redox) reactions, rearrangement of electrons among species occurs. In redox reactions, much more energy could be released than in acid–base reactions and safe handling of oxidants is very important. Which of the listed molecules is a strong oxidant?

a. H_2O
b. KNO_3
c. NaCl
d. CH_3COOH

Note: Correct answers are given in red. These need to be removed before the test is used.

TESTING PRIOR KNOWLEDGE

Select the correct answer.

1. Some substances are called electrolytes. We distinguish between strong and weak electrolytes. Which one on the list is a strong electrolyte in an aqueous solution?

a. HCOOH
b. CH_3COOH
c. H_2S
d. NaCl

2. Acids and bases belong to the class of electrolytes. Electrolytes conduct electric current. For this to happen, species which are called ions must be present in a solution. Which of the listed species is an ion?

a. H_2O
b. NO_3^-
c. CH_3COOH
g. CH_4

3. Which of the listed chemicals is the acid in an aqueous solution?

a. NaOH
b. $Ca(OH)_2$
c. HCl
d. NaCl

4. The pH of an acid solution is:

a. 7
b. higher than 7
c. lower than 7
d. 14

5. A buffer is a mixture (solution) of a weak acid and its salt with a strong base or a weak base and its salt with a strong acid. Buffers are used to keep the pH of a solution constant. Several buffers are in our body to keep the pH of blood constant at 7.4. Find one of the buffers in the list below.

a. acetic acid – sodium acetate solution
b. hydrochloric acid – sodium chloride solution
c. potassium bromide – sodium chloride solution
d. lithium fluoride – sulfuric(VI) acid

7. In oxidation–reduction (redox) reactions, rearrangement of electrons among species occurs. In redox reactions, much more energy could be released than in acid–base reactions and safe handling of oxidants is very important. Which of the listed molecules is a strong oxidant?

a. H_2O
b. KNO_3
c. NaCl
d. CH_3COOH