

PURDUE UNIVERSITY

STUDY GUIDE AND SAMPLE EXAMINATION FOR CHEMISTRY 11100

This Study Guide describes topics to be mastered prior to attempting the examination to establish credit in Purdue's Chemistry 11100. The material can be studied from many of the textbooks on the market. A list of several of these is given below.

IMPORTANT:

1. Read this material thoroughly if you contemplate trying for advanced placement (and extra credit which counts toward graduation).
2. Study all the material listed in the outline.
3. Work many practice problems.
4. When you feel prepared for it, take the sample examination.
5. Come to the actual examination rested and confident.

SUBJECT MATTER - A Brief Outline

The subject matter of any General Chemistry course is varied. Chemistry 11100 deals with metrication, scientific units and measurements, the classification of matter, chemical and physical changes, chemical formulas and equations, energy in chemical reactions, characteristics of reactions, atomic weights, stoichiometry, atomic structure from both historical and modern aspects, periodic relationships among the elements, chemical bonding, solubility and solutions, chemical nomenclature, and oxidation-reduction. Interspersed in the preceding topics is descriptive chemistry pertinent to the various topics, e.g., hydrogen and oxygen are studied as examples of real gases and water as a common liquid.

The course is taken primarily by students in the College of Agriculture and the College of Health and Human Sciences and is designed to meet their needs. It is also taken by smaller numbers of students in the College of Liberal Arts and the College of Technology. A knowledge of high-school chemistry is presumed, and the course goes far beyond the high school level. A weekly three-hour laboratory is an integral part of the course and includes measurement techniques, the use of the analytical balance, titration techniques and experiments with various elements and compounds.

The topics presented in the outline should be studied prior to attempting the sample examination included with this study guide. These topics are broken down under several headings according to those found in the texts currently in use at Purdue.

In preparing for the examination it is important to work many problems. However, the problems may test your recall of relations like $D = \text{mass/volume}$ or $M = \text{moles/Liter}$, and also should be used to measure your understanding of the concepts and principles involved.

At the end of this Study Guide you will find a sample examination over this material. Allow one hour for the exam. Naturally, it does not cover every point. No examination extending over a reasonable time period can do that.

SPECIAL NOTE:

A word of advice concerning the taking of the actual examination for credit. No one does well on an examination when he/she is excessively fatigued. Therefore, you are urged to provide yourself an adequate rest period before taking the actual examination. If your trip to the campus necessitates travel into the late hours of the night or an extremely early departure from your home, you may be well advised to allow for a one night's rest in the Lafayette area before taking the examination. Many students who are unsuccessful with the examination tell us that failing to take the above precautions contributed strongly to this result. Most such students find that their first year was somewhat less rewarding than it might have been because of the time spent retracing materials studied in high school. Please consult your advanced credit schedule for the actual time and place of the examination. It is usually given both morning and afternoon.

The texts referred to below may be ordered from any of the local book stores:

Local Book Stores:

University Book Store, 360 State Street, W. Lafayette, IN 47906.

Follett's Purdue Book Store, Purdue Service Center, W. Lafayette, IN 47906.

Texts:

Chemistry; BURDGE 4e, McGraw-Hill.

STUDY GUIDE FOR CHEMISTRY 11100

1. Fundamentals.

- The metric system.
- Basic SI units and prefixes, unit conversions.
- Significant figures - identify the number of significant figures in a number; rules for addition and subtraction, multiplication and division.

2. The mole, chemical formulas, and chemical equations.

Know the relationship between a mole of material and its weight, volume, concentration, or number of atoms or molecules.

Balance simple equations by inspection.

Given the formula for any compound, (a) name the elements that make up the compound; (b) indicate the number of each kind of atom represented; (c) determine the mass of each element in one mole of the compound; and (d) calculate the molecular (formula) weight.

Given a sample of a compound (or its mass) and its formula, calculate, (a) the number of moles of the compound in the sample; (b) the number of moles of any one element in the sample; (c) the number of molecules (formula units) in the sample; and (d) the number of atoms of any one element in the sample.

3. Chemical stoichiometry.

Given a sample of some compound and a way to find the mass of each element, find the percent (by weight) of each element.

Given all reactants and products of a reaction (with the correct formulas) and any one of the following pieces of information, calculate any other piece of information on the list.

- (a) the number of moles of one of the products or reactants
- (b) the mass of any product or reactant
- (c) the number of molecules of any product or reactant
- (d) the number of atoms or mass of any single element involved
- (e) concentration and volume of a solution of reactant or product

Know the definition of exothermic and endothermic and how these terms are related to chemical reactions.

4. Properties of liquids and solutions.

Describe what occurs when a solid changes to a liquid or a liquid changes to a gas using a microscopic model. (a) What differences are there in the arrangement of the atoms of molecules; (b) what happens to the average speed of the particles; (c) what changes occur in the distance between particles; (d) what temperature changes occur, why?

Now how to calculate density and what measurements are required.

Take a chemical off of the shelf and prepare a specified volume of solution of some given concentration. (e.g., How would you prepare 200 mL of 0.1 M NaCl?)

Calculate the moles (or mass) of solute in a given volume of solution of known

concentration. (e.g., How many grams of NaCl are in 50 ml of a 1.5 M solution?)

Use a solution of some known concentration to prepare a solution of a specified lower concentration by proper dilution. (e.g., Make 100 ml of 0.22 M HCl from a solution of 2.5 M HCl.)

5. Periodic chemistry.

By referring only to the Periodic Table you should be able to give the common valence for any element in Group IA, IIA, IB, IIB, IIIB, IVB, VB, VIB, and VIIB.

Write the formula (including the proper charge) for any of the following polyatomic ions: ammonium, chlorate, hydroxide, sulfate, sulfite, carbonate, nitrate, nitrite, and phosphate, phosphite.

Identify a period, row, group, column, and family on the Periodic Table.

Be able to describe and use atomic radius, ionic radius, and electronegativity in terms of period trends.

6. Electronic structure of atoms.

Use the relationship between the speed of light, frequency, and wavelength.

Given the symbol for an element or ion and a Periodic Table, specify the number of protons in the nucleus of that particle and the number of electrons around the nucleus.

Given an isotope of an element, find the number of protons, neutrons and electrons used.

From the position of an element in the Periodic Table, indicate the number of electrons in the outer or valence shell. Conversely, knowing the number of electrons in the valence shell, identify the chemical family (group) to which the element belongs.

Write electron configurations for elements and ions. Identify the number of valence e^- from the electron configuration.

7. Chemical bonding and related phenomena.

Given the symbol, name or atomic number for any element, write the Lewis structure to represent that element.

Write the Lewis (electron-dot) structure for any compound containing no more than three elements.

Identify the shapes of molecules; use bond polarity and shape to describe the polarity of a molecule.

From the chemical and physical properties of a compound (e.g., electrical conductivity, melting point, heat of fusion, etc.) predict whether the compound is ionic or covalent.

Use a table of electronegativities to predict whether the bond between two given elements will be ionic or covalent.

CHEMISTRY 11100 - SAMPLE EXAM

Choose the one best response for each question.

- _____ 1. When fluorine reacts with calcium to produce CaF_2 , the bond formed is:
- (a) ionic.
 - (b) covalent.
 - (c) metallic.
- _____ 2. A chemical reaction that absorbs heat from its surroundings is termed:
- (a) exothermic.
 - (b) isothermal.
 - (c) endothermic.
 - (d) caloric.
 - (e) spontaneous.
- _____ 3. In the sum of $2.82 + 1.096$ there are ____ significant figures.
- (a) 3
 - (b) 4
 - (c) 5
 - (d) 6
 - (e) 17
- _____ 4. What is the correct name for CCl_4 ?
- (a) carbon tetrachloride
 - (b) carbon chloride
 - (c) carbon(IV) chloride
 - (d) carbon chlorate
- _____ 5. $^{64}_{30}\text{Zn}$ and $^{68}_{30}\text{Zn}$ are examples of:
- (a) isotopes.
 - (b) allotropes.
 - (c) phenemones.
 - (d) cations.
 - (e) anions.
- _____ 6. Which one of the following combinations of names and formulas is **incorrect**?
- (a) H_3PO_4 phosphoric acid
 - (b) HNO_3 nitric acid
 - (c) NaHCO_3 sodium carbonate
 - (d) H_2CO_3 carbonic acid
 - (e) KOH potassium hydroxide

_____ 7. What ions and how many of each are formed for every one formula unit of $\text{Al}(\text{NO}_3)_3$ that dissociates in water?

- (a) One Al^{3+} ion and one NO_3^- ion
- (b) One Al^{3+} ion and three NO_3^- ions
- (c) One Al^{3+} ion, one N^{3-} ions, and three O^{2-} ions
- (d) One Al^{3+} ion, three N^{3-} ions, and nine O^{2-} ions

_____ 8. When NaCl dissolves in water, it is present as:

- (a) NaCl molecules.
- (b) Na atoms and Cl atoms.
- (c) Na^+ cations and Cl^- anions.
- (d) NaOH and HCl .
- (e) $(\text{NaCl})_2$ dimers.

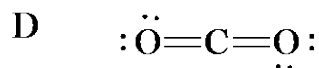
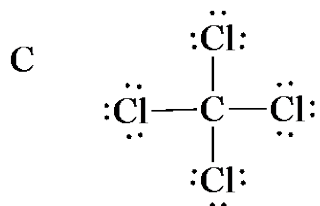
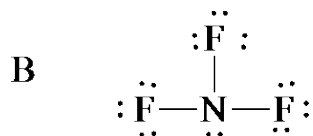
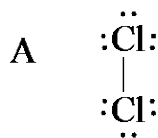
_____ 9. The name for the ion, SO_4^{2-} is:

- (a) sulfate.
- (b) sulfite.
- (c) sodium.
- (d) saccharin.
- (e) sulfide.

_____ 10. Arrange P , Si , and O in order of increasing atomic radius. (Smallest to largest.)

- (a) $\text{Si} < \text{O} < \text{P}$
- (b) $\text{P} < \text{Si} < \text{O}$
- (c) $\text{O} < \text{Si} < \text{P}$
- (d) $\text{O} < \text{P} < \text{Si}$

_____ 11. Which of the following molecules is polar?



_____ 12. The proper Lewis structure (electron-dot structure) for N_2 is:

- (a) $\text{:}\ddot{\text{N}}::\ddot{\text{N}}\text{:}$ (d) $\cdot\ddot{\text{N}}:\ddot{\text{N}}\cdot$
 (b) $\ddot{\text{N}}:\ddot{\text{N}}:$ (e) :N::N:
 (c) :N:::N:

13. Which of these compounds is a **not** a strong electrolyte when dissolved in water?

- (a) $\text{C}_6\text{H}_{12}\text{O}_6$
 (b) NaNO_3
 (c) HCl
 (d) KOH

_____ 14. A given liquid occupies 445 mL and weighs 540 g. The density of this liquid is (units missing):

- (a) 0.73 (b) 0.82 (c) 0.85 (d) 1.21 (e) 1.24

_____ 15. The ion whose outer shell electronic configuration is $2s^2 2p^6$ is:

- (a) Ne (b) Li^+ (c) F^- (d) S^{2-} (e) Ca^{2+}

_____ 16. The elements of Group VII are known as the:

- (a) Alkaline earth metals (d) Ideal gases.
 (b) Nobel gases. (e) Halogens.
 (c) Alkali metals.

_____ 17. Which of the following compounds has ionic bonds?

- (a) NaCl
 (b) O_2
 (c) CH_4
 (d) SiO_2

_____ 18. To what volume must you dilute 200.0 mL of a 6.00 M solution of NaCl to obtain a 0.30 M solution of NaCl ?

- (a) 1.0 L (b) 2.0 L (c) 4.0 L (d) 0.5 L (e) 7.8 L

_____ 19. Calculate the molar mass of $(\text{NH}_4)_3\text{AsO}_4$.

- (a) 417.80 g/mol
 (b) 193.03 g/mol
 (c) 165.02 g/mol

- (d) 156.96 g/mol
- (e) None of the above are close to the correct answer.

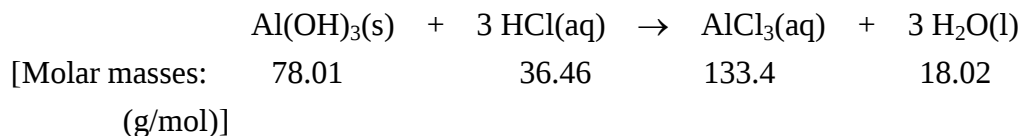
20. What is the molarity of a sodium chloride solution created from 6.30 g NaCl dissolved into 785 mL of solution? (NaCl 58.44 g/mol)

- (a) 1.25 M
- (b) 0.137 M
- (c) 8.03×10^{-3} M
- (d) 1.37×10^{-4} M

_____ 21. Convert 0.928 kg to mg.

- (a) 9.28×10^{-3} mg
- (b) 9.28×10^1 mg
- (c) 9.28×10^4 mg
- (d) 9.28×10^5 mg
- (e) 9.28×10^6 mg

22. Several brands of antacid tablets use aluminum hydroxide to neutralize excess acid according to the balanced chemical equation. What quantity of HCl, in grams, can a tablet with 0.750 g of $\text{Al}(\text{OH})_3$ consume .



- (f) 0.351 g
- (g) 1.05 g
- (h) 4.81 g
- (i) 27.35 g

23. 25.00 mL of 0.125 M aluminum chloride solution is mixed with 25.00 mL of 0.100 M sodium hydroxide solution. Calculate the mass of aluminum hydroxide that will be formed ($\text{Al}(\text{OH})_3$, 78.01 g/mol).

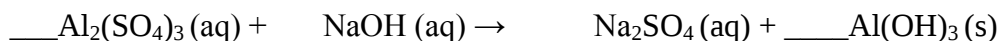
- (a) 0.065 g
- (b) 0.179 g
- (d) 0.309 g

(c) 0.244 g

_____ 24. Calculate the number of oxygen atoms in 29.34 g of sodium sulfate, Na_2SO_4 .

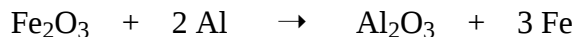
- (a) 1.244×10^{23} O atoms
- (b) 4.976×10^{23} O atoms
- (c) 2.915×10^{24} O atoms
- (d) 1.166×10^{25} O atoms
- (e) None of the above are close to the correct answer.

_____ 25. What are the coefficients of the following chemical equation when it is balanced?



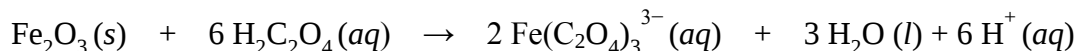
- (a) 1, 1, 1, 1
- (b) 1, 2, 1, 2
- (c) 1, 6, 3, 2
- (d) 1, 3, 3, 2

_____ 26. How much iron can be made from 495 g of aluminum by the following process:



- (a) 683 g
- (b) 743 g
- (c) 1020 g
- (d) 1540 g
- (e) 2310 g

_____ 27. Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) can remove rust (Fe_2O_3) caused by bathtub rings according to the balanced chemical equation below.



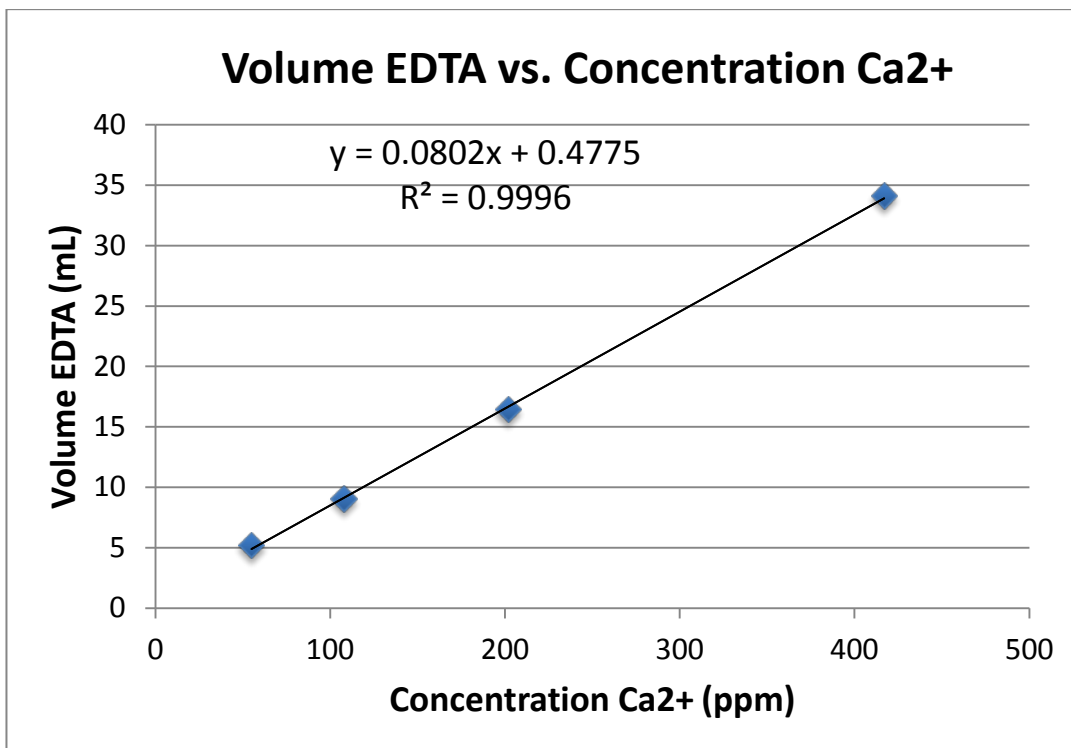
Calculate the number of GRAMS of rust that can be removed by 3.00×10^2 mL of a 0.600 M solution of oxalic acid. (Molar masses $\text{H}_2\text{C}_2\text{O}_4$ 90.036 g/mol; Fe_2O_3 159.70 g/mol)

- (a) 4.79 g
- (b) 28.7 g
- (c) 47.9 g
- (d) 4790 g

_____ 28. How many grams of water will be formed when a mixture containing 1.93 g C_2H_4 and 5.92 g O_2 is ignited in a closed vessel?

- | | |
|-------------|-------------|
| (a) 2.22 g | (d) 12.81 g |
| (b) 5.89 g | (e) 14.50 g |
| (c) 10.21 g | |

29. In the “How hard is “hard” water?” lab a calibration curve was created for four calcium standards with EDTA. The trendline is on the graph. If a sample of tap water required 21.55 mL of EDTA to titrate, what is the concentration of calcium in the water?



- (a) 1.69 ppm
(b) 220 ppm
(c) 263 ppm
(d) 275 ppm

*****Additional help (problem solving by topic) is available here:**
<http://www.chem.purdue.edu/gchelp/>

This is a site with multiple questions on many different topics!

ANSWERS TO 11100 SAMPLE EXAM

1. A
2. C
3. A
4. A
5. A
6. C
7. B
8. C
9. A
10. D
11. B
12. C
13. A
14. D
15. C
16. E
17. A
18. C
19. B
20. B
21. D
22. B
23. A
24. B
25. C
26. D
27. A
28. A
29. C