Practice Test Questions 11
Nonmetals Part 2 (Groups 14-17 and Boron)

1.
(a) Give the symbol <u>and</u> name for a halogen. _____ and ______________________

(b) Give the formula <u>and</u> name for an oxoacid. ______ and ____________________________

(c) Give the formula <u>and</u> name for an oxoanion. _______ and __________________________

(d) Give the chemical formula for the halogen that is a liquid at room temperature. ________

(e) Give the symbol for a chalcogen that is also a metalloid. _________

2. Complete the following table.

Formula	**Name**
$CrPO_3$	
	copper(II) nitrite
	iron(III) sulfate
HNO_3	
	copper(II) perchlorate
	lithium phosphate
	calcium hydrogen carbonate
	calcium nitrate
Na_2SO_3	
H_2SO_3	

3. Write a balanced chemical equation for each of the reactions described below. *Include states of matter*.

(a) Liquid phosphorus trichloride is prepared from white phosphorus and chlorine gas.

(b) Sulfur dioxide and water vapour react in the upper atmosphere

(c) Ammonia gas is bubbled into a solution of aqueous acid.

(d) Ammonia gas and boron trifluoride gas are mixed and a single solid product forms.

(e) Solid P_4O_{10} reacts with liquid H_2O.

(f) Ozone decays to oxygen.

4.

(a) Draw Lewis structures for each of the allotropes of oxygen.

(b) On your structures, clearly label any bond angles (redrawing the structure if necessary to show the proper shape) and identify the molecular geometry.

5. Name three allotropes of carbon and briefly describe each.

6. Does nitrogen exhibit allotropy? Does phosphorus? Describe the allotropes formed.

7. What is unusual about the bonding in diborane (B_2H_6)? Include a diagram in your answer.

8.

(a) Draw Lewis structures (and resonance structures, if appropriate) for the following ions:

$$IF_2^+ \quad \text{and} \quad IF_4^+$$

Clearly show the formal charge of any atom with a non-zero formal charge.

(b) Predict the geometry of each ion.

(c) Calculate the oxidation state of the iodine atom in each ion.

(d) Which of the two ions would you expect to be the stronger oxidizing agent and why?

9. Rank the following compounds from lowest pK_a to highest pK_a, and explain your logic.

$$H_3PO_4, H_2PO_4^-, H_2SO_4, H_2O$$

10.

(a) Draw Lewis structures for:

i. HCl ii. $HClO_2$ iii. $HClO_3$

(b) Determine the oxidation state of the chlorine atom in each of these acids.

i. HCl ii. $HClO_2$ iii. $HClO_3$

(c) The approximate pK_a of HCl is -7. Calculate the approximate pK_a values for $HClO_2$ and $HClO_3$.

i. $HClO_2$ ii. $HClO_3$

(d) Rank the three acids from strongest to weakest.

(e) The names for these acids are chloric acid, chlorous acid and hydrochloric acid. Which is which?

11.

(a) Give formulas for the following common acids: nitric acid, sulfuric acid, hydrobromic acid, perchloric acid, carbonic acid.

(b) What is the oxidation state of the central atom in each of these compounds?

(c) How does the oxidation state of each common acid relate to the periodic group number of the element?

12.

(a) Arrange the oxides below in order from the most acidic through amphoteric to the most basic: Al_2O_3, BaO, CO_2, Cl_2O_7, SO_3

(b) Assign oxidation states to the non-oxygen element in each oxide.

(c) What is the relationship between the oxidation state of the central atom(s) and the acidity of the oxide?

13. Xenic acid (H_2XeO_4) is an oxoacid.

(a) Draw a Lewis structure for H_2XeO_4.
Include any non-zero formal charges on the appropriate atoms.

(b) What is the electron group geometry of Xe in H_2XeO_4?

(c) What is the molecular geometry of Xe in H_2XeO_4?

(d) What is the oxidation state of the Xe atom in H_2XeO_4?

(e) Should H_2XeO_4 acid act as a reducing agent, an oxidizing agent or neither?

(f) Calculate approximate pK_a values for H_2XeO_4 and $HXeO_4^-$.

(g) Is H_2XeO_4 predicted to be a strong acid or a weak acid?

14. Describe how chlorine gas is produced industrially.

15. Describe the Haber-Bosch process for the industrial production of ammonia. Be sure to include balanced equations for any relevant reactions and explain why any special conditions (e.g. use of a catalyst, pressure, heat, etc.) are necessary.

16. There are multiple methods used for the industrial production of sulfur. In the southern United States, the Frasch process is the main method used. In western Canada, the Claus process is the main method used. Describe **one** of these methods, including any relevant balanced chemical equations. Make sure that you clearly indicate which process you are describing.

17. A chunk of white phosphorus weighing 6.58 grams is put in a 750 mL flask containing dry argon (which is then removed using a vacuum, leaving only the phosphorus in the flask).

A separate 750 mL flask contains 3.15 bar of fluorine gas (at 19.65 °C).

The two flasks are connected so that the two compounds can react, producing phosphorus trifluoride (a gas that is colourless and odourless, but highly toxic).

(a) Write a balanced chemical equation for this reaction.

(b) What mass of phosphorus trifluoride is produced in this reaction?

18. One way to make sodium sulfite ($Na_2SO_{3(aq)}$) is to bubble sulfur dioxide ($SO_{2(g)}$) through an aqueous sodium hydroxide solution.

(a) Write a balanced chemical equation for this reaction.

(b) You have 0.75 L of $SO_{2(g)}$ at 25 °C and 94 kPa. The largest reaction flask that you have available will only hold 0.100 L of aqueous solution. What is the minimum concentration of sodium hydroxide solution that you can use if you need 0.100 L of the solution to react with all of the $SO_{2(g)}$?

19.

(a) Phosphorus reacts readily with halogens. If 6.1 grams of white phosphorus reacts with chlorine gas to produce 41 grams of a phosphorus halide, what is the molecular formula of the product?

(b) Draw a Lewis structure of the product and indicate its molecular geometry.