University of Saskatchewan  
Department of Chemistry  
CHEMISTRY 112.3  
Final Examination  
April 9, 2012 (9:00 AM – 12:00 pm)  
EXAM COPY A1

Last Name (please print clearly!)____________________________________________

Given Name (please print clearly!)__________________________________________

Student Number________________________ Signature__________________________

Please indicate your section:

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<tr>
<td>02</td>
<td>MWF 9:30 AM</td>
<td>Dr. A. Grosvenor</td>
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<td>04</td>
<td>TR 10:00 AM</td>
<td>Dr. P. Ahiahonu</td>
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<td>North West</td>
<td>Dr. R. Catton</td>
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INSTRUCTIONS - PLEASE READ THIS FIRST!

1. This is a closed-book examination. A data-sheet with a Periodic Table is attached to the last page of this examination paper (you can remove this page from the booklet for convenience)

2. Simple scientific calculators are permitted. Alphanumeric calculators and those capable of storing equations are not permitted. Cell phones, personal digital assistants, electronic dictionaries, etc. are not allowed. No equations may be stored in memory of electronic devices.

3. This examination paper has 15 pages, including the data sheet. 1.5 marks per question awarded in Part A.

4. Answer the questions in Part A by circling the response on this paper AND by filling out the corresponding response on the blue opscan sheet USING A SOFT-LEAD PENCIL ONLY. No deductions will be made for incorrect answers. If you change your mind, erase the incorrect answer carefully from the blue optical scan sheet. In the event of a discrepancy, the response on the opscan sheet will count.

5. If you have not yet done so, fill in the top of this paper now and code your student number and indicate the exam copy code on the computer opscan sheet in soft HB lead pencil.

6. Answer the questions in Part B on this examination paper, showing all work.

7. HAND-IN ALL of your material (question sheet and computer sheet)

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Part A

1. What is the correct formula for potassium dihydrogen phosphate hexahydrate?
   A. KHPO₄·6H₂O
   B. K₂HPO₄·5H₂O
   C. KH₂PO₄·5H₂O
   D. K₂HPO₄·6H₂O
   E. None of the above

2. Which of the following represents a 0.100 mol/L KCl solution?
   A. 1000 mL of solution containing 74.6 g KCl
   B. 100 g KCl in 1.00 L of solution
   C. 5.00 L of solution containing 373 g KCl
   D. 7.46 mg KCl per mL of solution
   E. None of the above

3. Given the following unbalanced chemical reaction, what volume of 0.650 mol/L K₂CrO₄ is required to completely precipitate all of the silver ions in 415 mL of 0.186 mol/L AgNO₃ as Ag₂CrO₄(s).
   \[ \text{AgNO}_3 \text{(aq)} + \text{K}_2\text{CrO}_4 \text{(aq)} \rightarrow \text{Ag}_2\text{CrO}_4 \text{(s)} + \text{KNO}_3 \text{(aq)} \]
   A. 59.4 mL
   B. 119 mL
   C. 593 mL
   D. 29.7 mL
   E. 178 mL

4. An element has two naturally occurring isotopes. The first isotope has a mass of 78.9183 g/mol and an abundance of 54.54%, while the second isotope has a mass of 80.9163 g/mol. What is the average atomic mass of the element?
   A. 80.12 g/mol
   B. 79.83 g/mol
   C. 78.99 g/mol
   D. 79.13 g/mol
   E. 80.68 g/mol

5. Determine the mass of nitrogen that can be formed from the reaction of 50.0 g N₂O₄ and 45.0 g N₂H₄.
   Use the following balanced chemical reaction to answer the question:
   \[ \text{N}_2\text{O}_4 \text{(l)} + 2 \text{N}_2\text{H}_4 \text{(l)} \rightarrow 3\text{N}_2 \text{(g)} + 4\text{H}_2\text{O} \text{(g)} \]
   A. 13.3 g
   B. 45.6 g
   C. 59.0 g
   D. 45.0 g
   E. 105 g
6. The food flavor enhancer monosodium glutamate (MSG) has the composition 13.6% Na, 35.5% C, 4.8% H, 8.3% N and 37.8% O, by mass. What is the empirical formula of MSG?

A. C₃H₈NO₄  
B. NaC₅H₈NO₄  
C. NaC₃H₇NO₄  
D. Na₂C₅H₈NO₄  
E. NaC₅H₈NO₄

7. An 6.72 g sample of Ag₂O decomposes into solid silver and O₂(g). If the gas is collected over water at 25 °C and 0.986 atm of pressure, what is the volume of O₂(g) gas collected? (The vapour pressure of water at 25 °C is 0.0313 atm.)

The balanced chemical equation for the decomposition is: 2Ag₂O(s) → 4Ag(s) + O₂(g).

A. 327 mL  
B. 371 mL  
C. 350 mL  
D. 305 mL  
E. 343 mL

8. The empirical formula of a gaseous fluorocarbon is CF₂. Determine the molecular formula of this compound if 1.55 g of this compound occupies 0.174 L at STP.

A. CF₂  
B. C₂F₄  
C. C₃F₆  
D. C₄F₈  
E. C₅F₁₀

9. An adult takes about 15 breaths per minute, with each breath having a volume of 500 mL. If the air that is inhaled is “dry,” but the exhaled air at 1 atm pressure is saturated with water vapor at 37 °C (body temperature), what mass of water is lost from the body during a 3 hour period? (The vapour pressure of water at 37.0 °C is 0.0636 atm.)

A. 60.8 g  
B. 48.3 g  
C. 30.4 g  
D. 56.7 g  
E. 72.3 g
10. A gas mixture of Ar\(_g\), N\(_2\)\(_g\) and O\(_2\)\(_g\) has a total pressure of 0.850 atm and occupies a volume of 2.00 L when its temperature is 290 K. What is the mass of N\(_2\)\(_g\) present, if the partial pressure of Ar\(_g\) is 0.450 atm and the partial pressure of O\(_2\)\(_g\) is 0.210 atm?

A. 1.77 g  
B. 2.29 g  
C. 0.45 g  
D. 1.21 g  
E. 0.57 g

11. Which statement properly describes the differences between a real gas and an ideal gas?

A. The molar volume of a real gas is larger than predicted by the ideal gas law at high pressures.  
B. Intermolecular attractive forces make the real pressure less than the ideal gas law would predict at low temperatures.  
C. Because real molecules attract each other, the molar volume of a real gas is smaller than predicted by the ideal gas law at low temperatures.  
D. All of these statements are correct  
E. None of these statements are correct

12. Calculate the amount of heat (in kJ) required to raise the temperature of a 2.15 kg iron bar by 24.0 °C. (The specific heat of iron is 0.473 J/g °C.)

A. 109 kJ  
B. 51.6 kJ  
C. 23.6 kJ  
D. 24.4 kJ  
E. 11.3 kJ

13. Using Hess’s, law calculate ΔH°\(_{\text{rxn}}\) for the reaction: SO\(_2\)\(_g\) + NO\(_2\)\(_g\) \rightarrow SO\(_3\)\(_g\) + NO\(_g\)).

Given:

\[2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{SO}_3\text{(g)} \quad \Delta H^{\circ}_{\text{rxn}} = -197.8 \text{ kJ}\]
\[\text{O}_2\text{(g)} + 2\text{NO}_2\text{(g)} \rightarrow 2\text{NO}_2\text{(g)} \quad \Delta H^{\circ}_{\text{rxn}} = -114.14 \text{ kJ}\]

A. -83.66 kJ  
B. -311.9 kJ  
C. +155.9 kJ  
D. -155.9 kJ  
E. -41.83 kJ

14. What is the formation reaction for CdSO\(_4\)(s)?

A. Cd(NO\(_3\))\(_2\)(aq) + H\(_2\)SO\(_4\)(aq) \rightarrow 2\ HNO\(_3\)(aq) + CdSO\(_4\)(s)  
B. Cd\(_{2+}\)(aq) + 2SO\(_4\)^{2-}(aq) \rightarrow CdSO\(_4\)(s)  
C. Cd\(_s\) + S\(_s\) + 2 O\(_2\)(g) \rightarrow CdSO\(_4\)(s)  
D. CdO\(_s\) + SO\(_3\)(l) \rightarrow CdSO\(_4\)(s)  
E. All of these answers are correct
15. Determine the change in enthalpy for the following equation:
\( \text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{HCl}(\text{g}) + 1/2 \text{O}_2(\text{g}) \)

Given information:
\[ \Delta H_{\text{f}}^\circ \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{O}_2(\text{g}) = 0.00 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{O}(\text{g}) = 249.2 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{HCl}(\text{g}) = -92.3 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{Cl}_2(\text{g}) = 0.00 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{HCl}(\text{aq}) = -167.2 \text{ kJ/mol} \]
\[ \Delta H_{\text{f}}^\circ \text{Cl}(\text{g}) = 121.3 \text{ kJ/mol} \]

A. -104.5 kJ  
B. -470.4 kJ  
C. 89.3 kJ  
D. 101.2 kJ  
E. -92.6 kJ

16. What is the change in internal energy (\( \Delta E \)) of a system if it performs 356 J of work and loses 289 J of heat?

A. +645 J  
B. -645 J  
C. +67 J  
D. -67 J  
E. +134 J

17. What is the frequency of radiation with a wavelength of 440 nm.

A. \( 6.82 \times 10^{14} \text{ s}^{-1} \)  
B. \( 6.81 \times 10^5 \text{ s}^{-1} \)  
C. \( 1.51 \times 10^{-27} \text{ s}^{-1} \)  
D. \( 4.95 \times 10^{-12} \text{ s}^{-1} \)  
E. \( 2.02 \times 10^{-11} \text{ s}^{-1} \)

18. Determine the wavelength of a single photon if the total energy of a mol of photons is 200.0 kJ/mol.

A. 167 nm  
B. 598 nm  
C. \( 5.98 \times 10^5 \text{ nm} \)  
D. \( 1.67 \times 10^5 \text{ nm} \)  
E. 242 nm

19. Which one of the following set of quantum numbers is not allowed?

A. \( n = 3, \ l=2, \ m_l=1 \)  
B. \( n = 3, \ l=0, \ m_l=0 \)  
C. \( n = 3, \ l=3, \ m_l=1 \)  
D. \( n = 3, \ l=2, \ m_l=-1 \)  
E. \( n = 3, \ l=1, \ m_l=1 \)

20. Which of the following orbitals will have the lowest (most negative) energy if we are considering the
21. Which set of quantum numbers would best describe the highest energy electron of nickel (Ni) metal?

A. \( n = 3, \ l=2, \ m_l=0, m_s=\frac{1}{2} \)
B. \( n = 4, \ l=2, \ m_l=0, m_s=\frac{1}{2} \)
C. \( n = 3, \ l=1, \ m_l=0, m_s=\frac{1}{2} \)
D. \( n = 3, \ l=2, \ m_l=0, m_s=0 \)
E. \( n = 3, \ l=2, \ m_l=\frac{1}{2}, \ m_s=\frac{1}{2} \)

22. What is the effective nuclear charge \( (Z_{eff}) \) that acts on the valence electrons of antimony (Sb)?

A. +15
B. +3
C. +5
D. +2
E. +13

23. Rank the following elements from lowest-to-highest first ionization energy: P, Al, Mg, Na.

A. Na<Mg<Al<P
B. P<Mg<Al<Na
C. Al<Na<Mg<P
D. P<Al<Mg<Na
E. Na<Al<Mg<P

24. Place the following in order of decreasing radius: \( \text{S}^{2-}, \text{Ar}, \text{Ca}^{2+}, \text{Rb} \).

A. (largest) Ar>Rb>Ca^{2+}>S^{2-} (smallest)
B. (largest) Rb>S^{2}->Ar>Ca^{2+} (smallest)
C. (largest) S^{2}->Ar>Ca^{2+}>Rb (smallest)
D. (largest) Rb>Ca^{2+}->Ar>S^{2} (smallest)
E. (largest) Ca^{2+}>S^{2}->Rb>Ar (smallest)
25. If an element had the following ionization energies, which column of the periodic table would it be found in?

\[ \text{IE}_1 = 738 \text{ kJ/mol, IE}_2 = 1450 \text{ kJ/mol, IE}_3 = 7730 \text{ kJ/mol, IE}_4 = 11600 \text{ kJ/mol} \]

A. 1A  
B. 2A  
C. 3A  
D. 4A  
E. 5A

26. Which cation, having a 2+ oxidation state, has the following electron configuration: \([\text{Xe}] 4f^{14} 5d^6\)?

A. \(\text{Tc}^{2+}\)  
B. \(\text{W}^{2+}\)  
C. \(\text{Eu}^{2+}\)  
D. \(\text{Os}^{2+}\)  
E. \(\text{Ba}^{2+}\)

27. Rank the bonds in the following binary compounds from most ionic to most covalent: \(\text{MgCl}_2, \text{MnO, KF, K}_2\text{O}\) (\(\chi_{\text{Mg}} = 1.2; \chi_{\text{Mn}} = 1.5; \chi_{\text{K}} = 0.8; \chi_{\text{Cl}} = 3.0; \chi_{\text{O}} = 3.5; \chi_{\text{F}} = 4.0\)).

A. (Most ionic) \(\text{MgCl}_2, \text{MnO, KF, K}_2\text{O}\) (Most covalent)  
B. (Most ionic) \(\text{KF, K}_2\text{O, MnO, MgCl}_2\) (Most covalent)  
C. (Most ionic) \(\text{KF, MgCl}_2, \text{K}_2\text{O, MnO}\) (Most covalent)  
D. (Most ionic) \(\text{MgCl}_2, \text{MnO, K}_2\text{O, KF}\) (Most covalent)  
E. (Most ionic) \(\text{MnO, MgCl}_2, \text{K}_2\text{O, KF}\) (Most covalent)

28. How many pairs of bonding and nonbonding electrons does Se have in \(\text{SeF}_4\)?

A. Bonding pairs = 1; Nonbonding pairs = 4  
B. Bonding pairs = 2; Nonbonding pairs = 3  
C. Bonding pairs = 3; Nonbonding pairs = 2  
D. Bonding pairs = 4; Nonbonding pairs = 1  
E. Bonding pairs = 5; Nonbonding pairs = 0

29. Which of the following can destabilize a bond between two atoms (i.e., increase its potential energy)?

A. Electron – Electron repulsion  
B. Electron – Nucleus attraction  
C. Nucleus – Nucleus repulsion  
D. A. and C.  
E. B. and C.
30. Select the correct Lewis structure for SrF$_2$, which is an ionic compound.

A. 
\[
\begin{array}{c}
\text{F} \\
\text{Sr} \\
\text{F}
\end{array}
\]

B. 
\[
\begin{array}{c}
\text{F} \\
\text{Sr} \\
\text{F}
\end{array}
\]

C. 
\[
\begin{array}{c}
\text{F} \\
\text{Sr} \\
\text{F}
\end{array}
\]

D. 
\[
\begin{array}{c}
\text{F} \\
\text{Sr} \\
\text{F}
\end{array}
\]

E. None of the above

31. What is the formal charge of N in NO$_3^{-}$?

A. 2-
B. 1-
C. 0
D. 1+
E. 2+

32. Which bond will have the lowest bond energy: C=C, C-C, C-N, C-S?

A. C=C
B. C-C
C. C-N
D. C-S
E. All of the bonds have the same energy

33. Which of the following has a trigonal planar electron geometry: AsF$_5$, BCl$_3$, ClF$_3$, XeF$_2$?

A. AsF$_5$
B. BCl$_3$
C. ClF$_3$
D. XeF$_2$
E. None of the above

34. Which of the following molecules does not have a net dipole moment (i.e., is non-polar): CH$_4$, NH$_3$, SO$_2$?

A. CH$_4$
B. NH$_3$
C. SO$_2$
D. All of the molecules listed are polar
E. None of the molecules listed are polar
35. What is the electron and molecular geometry of KrF₄?

A. Electron geometry = octahedral; Molecular geometry = square planar
B. Electron geometry = tetrahedral; Molecular geometry = tetrahedral
C. Electron geometry = trigonal bipyramidal; Molecular geometry = T-shaped
D. Electron geometry = octahedral; Molecular geometry = tetrahedral
E. Electron geometry = tetrahedral; Molecular geometry = trigonal pyramidal

36. Which intermolecular force is typically the weakest?

A. Dispersion forces
B. Dipole-Dipole forces
C. Hydrogen bonding
D. Ion-Dipole forces
E. All intermolecular forces have the same strength

37. Rank the following molecules in order of increasing strength of dipole-dipole forces acting between the molecules: CH₄, CH₂O, CH₃F.

A. CH₂O < CH₄ < CH₃F
B. CH₃F < CH₂O < CH₄
C. CH₄ < CH₂O < CH₃F
D. CH₄ < CH₃F < CH₂O
E. All of the molecules exhibit the same dipole-dipole strength

38. Rank the following molecules from lowest to highest boiling point: CH₄, He, HF, C₄H₁₀.

A. C₄H₁₀ < He < HF < CH₄
B. HF < He < C₄H₁₀ < CH₄
C. HF < C₄H₁₀ < CH₄ < He
D. CH₄ < HF < He < C₄H₁₀
E. He < CH₄ < C₄H₁₀ < HF

39. Which of the following will result in an increase in the viscosity of a fluid.

A. Raising the temperature
B. Increasing the strength of intermolecular forces between molecules
C. Decreasing the atmospheric pressure
D. Decreasing the surface tension
E. None of the above

40. What property affects the rate of vaporization of a substance?

A. The strength of intermolecular forces
B. Temperature
C. Surface area
D. All of the above
E. None of the above
Part B

Use the available space to answer the following questions.
Full marks will not be given if only the answer is provided (i.e., show steps).

1. a) Identify the following reactions as being either precipitation, acid-base neutralization, or redox reactions.
   If the reaction is a precipitation or acid-base neutralization reaction, provide the balanced net ionic equation.
   If the reaction is a redox reaction, identify the oxidizing agent and the reducing agent.
   (7 marks)

   a) \( \text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)} \)
   b) \( \text{K}_2\text{CO}_3(aq) + \text{NiCl}_2(aq) \rightarrow 2\text{KCl}(aq) + \text{NiCO}_3(s) \)
   c) \( \text{K}_2\text{CrO}_4(aq) + 2\text{AgNO}_3(aq) \rightarrow \text{Ag}_2\text{CrO}_4(s) + 2\text{KNO}_3(aq) \)
   d) \( \text{MnO}_2(s) + 4\text{H}^+(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O(l)} + \text{Cl}_2(g) \)
   e) \( \text{UO}^{2+}(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) + \text{H}^+(aq) \rightarrow \text{UO}_2^{2+}(aq) + \text{Cr}^{3+}(aq) + \text{H}_2\text{O(l)} \)
   f) \( \text{HCl}(aq) + \text{KOH}(aq) \rightarrow \text{KCl}(aq) + \text{H}_2\text{O(l)} \)
1. b) Using the following balanced chemical equation, determine the theoretical yield and the percent yield if 21.8 g of K₂CO₃ is produced from reacting 27.9 g of KO₂ with 29.0 L of CO₂ at STP: \(4\text{KO}_2(s) + 2\text{CO}_2(g) \rightarrow 2\text{K}_2\text{CO}_3(s) + 3\text{O}_2(g)\). (4 marks)

2. a) When 0.659 g of an organic compound (molar mass = 312 g/mol) undergoes combustion in a bomb calorimeter, the temperature rises from 298.2 K to 303.2 K. Find \(\Delta E_{\text{rxn}}\) for the combustion of this organic compound in kJ/mol. (The heat capacity of the bomb calorimeter is known to be 6.21 kJ/K.) (4 marks)
2. b) When 1 mol of a gas is burned at constant pressure, it produces 3540 J of heat and does 6.2 J of work on the surroundings. What are the numerical values for $\Delta E_{\text{constant, } P}$, $\Delta H_{\text{rxn}}$, $w$ and $q$ for this process? (3 marks)

2. c) Consider the following generic reaction: $2A + B \rightarrow C + 4D; \Delta H = 342.4 \text{ kJ}$. Determine the value of $\Delta H$ for the following reaction: $2D + \frac{1}{2} C \rightarrow A + \frac{1}{2} B$. (2 marks)

3. a) Using the Bohr model of the atom, calculate the energy of a photon produced when an electron transitions from $n_{2, \text{initial}} = 6$ to $n_{1, \text{final}} = 2$. (Hint: $E_n = -R_H \left( \frac{1}{n^2} \right)$) (2 marks)
3. b) Suppose that the photon produced in part 3.a) was absorbed by a metal atom, which resulted in the emission of a photoelectron. Using the equation for the photoelectric effect, what would the kinetic energy (KE) of the photoelectron be if its binding energy (Φ) was $1.000 \times 10^{-19}$ J? (1 mark)

3. c) Using the concepts of penetration and radial distribution functions, draw a diagram that shows why the 3s orbital is lower in energy (E) than the 3p orbitals, which are lower in energy than the 3d orbitals (i.e., $E_{3s} < E_{3p} < E_{3d}$). (2 marks)

3. d) Fill in the orbital diagram below so that it represents the proper ground-state electron configuration for S. (2 marks)

```
  1s  2s  2p  3s  3p  4s
```
4. Fill in the blanks in this table. The central atom for each species has been underlined.

(13 marks; 1 mark for each blank)

<table>
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<th>Species</th>
<th>Lewis Dot Structure</th>
<th>Resonance structures possible? (Yes/No)</th>
<th>Formal charge of the central atom</th>
<th>Electron group geometry of central atom</th>
<th>Molecular group geometry (shape)</th>
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### Periodic Table of the Elements

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### Physical Constants

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<td>Avogadro number</td>
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<td>Planck’s constant</td>
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<td>Gas Constant</td>
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<td>Molar volume of an ideal gas at STP</td>
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<td>Electron mass</td>
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<td>Speed of light in a vacuum</td>
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### Some SI Derived Units

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**STP Conditions:** $P = 1$ atm, Temperature $= 0$ °C = 273.15 K

*de Broglie's wavelength of particles: $\lambda = \frac{\hbar}{mv}$, Photoelectric effect: $KE = h\nu - \phi$*