

Name: \_\_\_\_\_

Student number: \_\_\_\_\_

**Chemistry 1A03**

**VERSION 1**

**Nov 9, 2018**

**McMaster University**

**Test 2**

Instructors: A. Chibba, L. Davis, D. Emslie, S. Greenberg

Duration: 120 minutes

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This test contains 21 numbered pages printed on both sides. There are **30** multiple-choice questions appearing on pages numbered 3 to 17. Pages 18 and 19 are extra space for rough work. Page 20 includes some useful data and equations, and there is a periodic table on page 21. You may tear off the last page to view the periodic table and the data provided.

**You must enter your name and student number on this question sheet, as well as on the answer sheet.** Your invigilator will be checking your student card for identification.

**You are responsible** for ensuring that your copy of the question paper is complete. Bring any discrepancy to the attention of your invigilator.

All questions are worth 1 mark - the total marks available are 30. There is **no** penalty for incorrect answers.

**BE SURE TO ENTER THE CORRECT VERSION OF YOUR TEST (shown near the top of page 1), IN THE SPACE PROVIDED ON THE ANSWER SHEET.**

**ANSWER ALL QUESTIONS ON THE ANSWER SHEET, IN PENCIL.**

Instructions for entering multiple-choice answers are given on page 2.

**SELECT ONE AND ONLY ONE ANSWER FOR EACH QUESTION** from the answers (A) through (E). **No work written on the question sheets will be marked.** The question sheets may be collected and reviewed in cases of suspected academic dishonesty.

Academic dishonesty may include, among other actions, communication of any kind (verbal, visual, *etc.*) between students, sharing of materials between students, copying or looking at other students' work. If you have a problem please ask the invigilator to deal with it for you. Do not make contact with other students directly. Try to keep your eyes on your own paper – looking around the room may be interpreted as an attempt to copy.

**Only Casio FX 991 MS or MS+ electronic calculators** may be used. They must NOT be transferred between students. Use of any aids other than those provided, is not allowed.

1. Which one of the following equilibrium constant expressions is **incorrect** based on the equilibrium reaction shown?

Equilibrium reaction	Equilibrium constant
A) $\text{H}_2(\text{g}) + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{Cu}(\text{s})$	$K = \frac{[\text{H}^+]^2}{P(\text{H}_2)[\text{Cu}^{2+}]}$
<b>B) <math>\text{S}(\text{aq}) \rightleftharpoons \text{S}(\text{org})</math></b>	<b><math>K = \frac{[\text{S}(\text{aq})]}{[\text{S}(\text{org})]}</math></b>
C) $\text{NH}_3(\text{g}) + \text{HNO}_3(\text{g}) \rightleftharpoons \text{NH}_4\text{NO}_3(\text{s})$	$K = \frac{1}{P(\text{NH}_3)P(\text{HNO}_3)}$
D) $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$	$K = P(\text{CO}_2)$

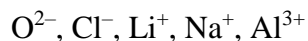
- For answer B, the equilibrium constant, K, should be equal to  $[\text{S}(\text{org})]/[\text{S}(\text{aq})]$

2. When comparing the elements **Al, Si, P, S, and Cl**, which one of the following statements is **false**?

- A) The order of increasing ionization energy is  $\text{Al} < \text{Si} < \text{S} < \text{P} < \text{Cl}$   
 B) The order of increasing  $Z_{\text{eff}}$  for an outermost electron is  $\text{Al} < \text{Si} < \text{P} < \text{S} < \text{Cl}$   
 C) **The order of increasing magnitude of electron affinity is  $\text{Al} < \text{Si} < \text{S} < \text{P} < \text{Cl}$**   
 D) The order of increasing atomic radius is  $\text{Cl} < \text{S} < \text{P} < \text{Si} < \text{Al}$

- Answer A is correct: the *general* trend is for ionization energy (IE) to increase across a period ( $\text{Al} < \text{Si} < \text{S} < \text{Cl}$ ), but P has a half-filled p-shell, which gives it a higher IE than S.
- Answer B is correct:  $Z_{\text{eff}}$  increases across a period
- For answer C: the *general* trend for the electron affinity (EA) is that the EA becomes more negative across a period (this trend holds for groups 13, 14, 16 and 17). However, P has a half filled p-shell, so has a much less negative electron affinity than expected. Therefore, the *magnitude* of EA for P will not be greater than that of S; it will be less than that of Si.
- Answer D is correct: atomic radius decreases across a period

3. When arranging the species below from largest to smallest radius, which one is in the **middle**?



- A)  $\text{Li}^+$   
**B)  $\text{Na}^+$**   
C)  $\text{Cl}^-$   
D)  $\text{Al}^{3+}$

- The anions  $\text{O}^{2-}$  and  $\text{Cl}^-$  will be the largest.
- $\text{Li}^+$  will be smaller than  $\text{Na}^+$  (same group, different period)
- $\text{Al}^{3+}$  will be smaller than  $\text{Na}^+$  (the resulting ions both have a [Ne] configuration, but with different charges)
- Therefore,  $\text{Al}^{3+}$  and  $\text{Li}^+ < \text{Na}^+ < \text{O}^{2-}$  and  $\text{Cl}^-$

4. An orbital has the following quantum numbers:

$$n = 6, \ell = ?, m_\ell = 4$$

What could be the missing value(s) for the  $\ell$  quantum number?

- A) 5 only  
B) 0, 1, 2, 3, 4, or 5  
**C) 4 or 5**  
D) 4, 5, or 6

- If  $n = 6$ ,  $\ell$  can have values of 0, 1, 2, 3... $n-1$  (i.e. values of 0, 1, 2, 3, 4 and 5)
- $m_\ell$  can have values of  $-\ell, \dots, 0, \dots, +\ell$ . Therefore, only orbitals with  $\ell \geq 4$  can have  $m_\ell = 4$ , so the answer is C.

5. The empirical formula of a compound is  $\text{CH}_2$ . At  $125^\circ\text{C}$ , 0.445 g of this compound in the gas phase occupies a volume of 350. mL at a pressure of 1.00 bar. **What is the molecular formula** of this compound?

- A)  $\text{C}_4\text{H}_8$   
**B)  $\text{C}_3\text{H}_6$**   
C)  $\text{C}_2\text{H}_4$   
D)  $\text{C}_5\text{H}_{10}$

- Use  $PV = nRT$  to find  $n$ , and then since you know  $n$  and  $m$ , you can find  $M$  ( $M = m/n$ ).
- Alternatively, calculate the density of the gas ( $d = m / V = 0.445\text{g} / 0.350\text{L} = 1.27_{14} \text{ g L}^{-1}$ ), and then use  $PM = dRT$  to find  $M$  directly:  
$$M = dRT/P = 1.2714 \text{ g L}^{-1} \times 0.083145 \text{ L bar K}^{-1} \text{ mol}^{-1} \times 398.15 \text{ K} / 1 \text{ bar} = 42.09 \text{ g mol}^{-1}$$
- Formula =  $(\text{CH}_2)_n$ , and each  $\text{CH}_2$  unit has a molecular mass of approximately 14.  
 $n = 42.09 \text{ g mol}^{-1} / 14 \text{ g mol}^{-1} = 3 \rightarrow \text{Formula} = \text{C}_3\text{H}_6$

6. An electron in a hydrogen atom undergoes a transition from  $n = 4$  to  $n = 2$ . **How many** of the following statements are **true**?

- I. The transition results in emission of a photon.  
II. The atom resulting from this transition is in an excited state.  
III. The frequency of light involved in this transition is lower than the frequency of light for the transition from  $n = 3$  to  $n = 1$ .

- A) 1  
B) 2  
C) 0  
**D) 3**

- I. Correct (a photon is emitted when an electron falls from a higher to a lower energy level).  
II. Correct (the electron in a ground state hydrogen atom has  $n = 1$ ).  
III. Correct (the energy difference between adjacent energy levels decreases with increasing values of  $n$ , and frequency is proportional to energy).

7. What is the **molecular geometry** of  $\text{IF}_5$ ?

- A) See-saw
- B) Trigonal bipyramidal
- C) Square pyramidal**
- D) Octahedral

- $\text{IF}_5$  has one lone pair on iodine, leading to an octahedral electron pair geometry and a square pyramidal molecular geometry.

8. A stable, diamagnetic molecule,  $\text{EF}_3$ , has a **T-shaped** molecular geometry. Which one of the following could be the **identity of element E**?

- A) P
- B) Te
- C) F
- D) Br**

- If  $\text{EF}_3$  has a T-shaped molecular geometry, it must have two lone pairs, and in order for this to be a neutral compound, E must be from group 17 (because if you start with 7 valence electrons and use 3 of those electrons to form single bonds with F, there will be 4 non-bonding electrons left over to form 2 lone pairs).
- The identity of E is therefore either F or Br, and it can't be F because: (1) the central atom in T-shaped  $\text{EF}_3$  has 10 valence electrons and fluorine cannot have  $>8$  electrons, and (2)  $\text{F}_4$  is not a known allotrope of fluorine.

9. Which one of the following molecules is **non-polar**?

Molecule	Molecular geometry
<b>A) <math>\text{XeF}_4</math></b>	<b>square planar</b>
B) $\text{CF}_2\text{Cl}_2$	tetrahedral
C) $\text{SF}_5\text{Cl}$	octahedral
D) $\text{SF}_4$	see-saw

- Only  $\text{XeF}_4$  has no molecular dipole.

10. Which one of the following statements about a  $\text{BH}_2^-$  anion is **false**?

- A) The  $\text{BH}_2^-$  anion contains only one lone pair
- B) The H–B–H angle is slightly greater than  $120^\circ$**
- C) Boron in  $\text{BH}_2^-$  has less than an octet of valence electrons
- D) The  $\text{BH}_2^-$  anion has a negative formal charge on boron

- The  $\text{BH}_2^-$  anion is an  $\text{AX}_2\text{E}$  molecule with a trigonal planar electron pair geometry and a bent molecular geometry.
- The anion has only one lone pair, which is located on boron.
- The H–B–H angle will be **less** than the ideal  $120^\circ$  due to repulsion from the lone pair.
- Boron has only 6 valence electrons.
- Boron has a negative formal charge.

11. Which one of the following compounds has bonds with the **most covalent character**?

- A) ClF**
- B)  $\text{AlF}_3$
- C)  $\text{MgF}_2$
- D)  $\text{BF}_3$

- The question can be viewed as asking which compound has the smallest difference in electronegativity between the central atom and fluorine, and the answer is ClF.

12. Which one of the following statements about the  $\text{NO}^+$  ion is **false**?

- A) The formal charge on the nitrogen atom in both  $\text{NO}^+$  and nitrate anion is the same**
- B) The  $\text{NO}^+$  ion has two lone pairs
- C)  $\text{NO}^+$  is a polar molecule
- D) The average N–O bond length is shorter in  $\text{NO}^+$  than in a nitrate anion

- $\text{N}\equiv\text{O}^+$  has a triple bond between N and O, with a positive formal charge on oxygen.
- It has one lone pair on nitrogen and one on oxygen (2 in total).
- It is a polar molecule (since it involves bonding between two different elements; also, it has a formal positive charge on one of them).
- $\text{NO}_3^-$  has a formal positive charge on nitrogen (this is not analogous to the situation in  $\text{NO}^+$ , which does not have a formal charge on nitrogen).
- $\text{NO}_3^-$  has a double bond to one oxygen, and single bonds to the two remaining oxygen atoms, each one bearing a formal negative charge. The average N–O bond order in  $\text{NO}_3^-$  is 1.33, compared with 3.00 in  $\text{NO}^+$ , so the N–O bond will be stronger and shorter in  $\text{NO}^+$ .

13. Which one of the following statements about  $\text{SO}_4^{2-}$ ,  $\text{HSO}_4^-$  and  $\text{HSO}_3^-$  is **false**?

- A) The electron pair geometry for  $\text{HSO}_3^-$  is the same as that of  $\text{HSO}_4^-$   
 B) There are 9 lone pairs in  $\text{HSO}_4^-$   
 C) The average formal charge on the terminal oxygen atoms is more negative for  $\text{HSO}_3^-$  than  $\text{HSO}_4^-$   
**D) The average bond order for the terminal sulphur–oxygen bonds is higher for  $\text{SO}_4^{2-}$  than  $\text{HSO}_3^-$**

	10 lone pairs	9 lone pairs	8 lone pairs
Electron Pair Geom	tetrahedral	tetrahedral	tetrahedral
Ave. Formal Charge (on terminal O)	$-2/4 = -1/2$	$-1/3$	$-1/2$
Ave. terminal S-O Bond Order	$6/4 = 1.5$	$5/3 = 1.67$	$3/2 = 1.5$

14. Determine the pH of a saturated solution of iron(II) hydroxide ( $K_{\text{sp}} = 8.0 \times 10^{-16}$ ).

- A) 8.98  
 B) 7.85  
**C) 9.07**  
 D) 7.53

	$\text{Fe}(\text{OH})_2 (\text{s}) \rightleftharpoons \text{Fe}^{2+} (\text{aq}) + 2 \text{OH}^- (\text{aq})$		
I	---	0	0
C	---	+x	+2x
E	---	x	2x

$$K_{\text{sp}} = [\text{Fe}^{2+}(\text{aq})][\text{OH}^-(\text{aq})]^2 = x (2x)^2 = 4x^3$$

$$x = (K_{\text{sp}}/4)^{1/3} = 5.848 \times 10^{-6}$$

$$2x = [\text{OH}^-(\text{aq})] = 1.1696 \times 10^{-5}$$

$$\text{pOH} = -\log[\text{OH}^-(\text{aq})] = 4.93, \text{ and } \text{pH} = 14 - \text{pOH} = 9.07$$

15. Determine the **molar solubility** of strontium fluoride ( $K_{sp} = 2.5 \times 10^{-9}$ ) in a 0.10 M solution of  $\text{Sr}(\text{NO}_3)_2$ .

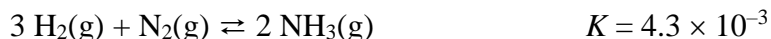
- A)  $8.6 \times 10^{-4}$   
 B)  $1.8 \times 10^{-4}$   
**C)  $7.9 \times 10^{-5}$**   
 D)  $5.0 \times 10^{-5}$

	$\text{SrF}_2 (\text{s})$	$\rightleftharpoons$	$\text{Sr}^{2+} (\text{aq})$	+	$2 \text{F}^- (\text{aq})$
I	---		0.1		0
C	---		+x		+2x
E	---		0.1+x		2x

$K_{sp} = [\text{Sr}^{2+}(\text{aq})][\text{F}^-(\text{aq})]^2 = (0.1+x)(2x)^2$ , but we can assume that x will be much smaller than 0.1, so  $K_{sp}$  simplifies to  $(0.1)(2x)^2 = 0.4x^2$

$x = (K_{sp}/0.4)^{1/2} = 7.91 \times 10^{-5} \text{ mol L}^{-1} \rightarrow$  This is the molar solubility of strontium fluoride because there is a 1:1 relationship between the amount of  $\text{SrF}_2$  that dissolves in 1L (molar solubility) and the concentration of  $\text{Sr}^{2+}$  (moles per litre) in a saturated solution of  $\text{SrF}_2$ .

16. For the synthesis of ammonia at 573 K:



A reaction mixture contains 19.0 bar  $\text{NH}_3$ , 28.6 bar  $\text{H}_2$  and 9.53 bar  $\text{N}_2$  in a 5.0 L flask at 573 K. Which statement is **true**?

- A)  $Q > K$  and the reaction shifts to the left  
 B)  $Q > K$  and the reaction shifts to the right  
 C)  $Q < K$  and the reaction shifts to the left  
**D)  $Q < K$  and the reaction shifts to the right**

$$Q = p\text{NH}_3^2 / (p\text{N}_2 \times p\text{H}_2^3) = (19)^2 / \{9.53 \times (28.6)^3\} = 1.62 \times 10^{-3}$$

Therefore,  $Q < K$ , and the reaction will shift to the right.



17. The reaction  $3 \text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$  is exothermic, with  $K = 4.3 \times 10^{-3}$  at 573 K. A vessel contains  $\text{N}_2$ ,  $\text{H}_2$  and  $\text{NH}_3$  at equilibrium at 573 K. Which one of the following changes will cause the **equilibrium constant to increase**?

- A) Adding more  $\text{NH}_3$  to the vessel.  
 B) Adding more  $\text{N}_2$  to the vessel.  
**C) Decreasing the temperature of the vessel.**  
 D) Decreasing the volume of the vessel.

- Equilibrium constants do not change upon addition of products or reagents (these changes shift reactions further to the left or to the right so that  $K$  remains the same).
- Equilibrium constants also do not change if the overall pressure is changed by decreasing or increasing the volume of the vessel.
- By contrast, changes in temperature can cause changes in the equilibrium constant. To determine whether the equilibrium constant will increase or decrease, consider heat in an exothermic reaction to be a product, and then apply Le Chatelier's principle. If heat is a product, increasing the temperature (adding heat) will push the reaction to the left hand side. Conversely, decreasing the temperature will push the reaction to the right hand side, giving more products and less reactants, increasing  $K$  (answer C).

18. Consider the reaction:  $\text{A}_2(\text{g}) \rightleftharpoons 2 \text{A}(\text{g})$ .

A flask is charged with 1.50 bar  $\text{A}_2$  and 1.80 bar A and maintained at a constant temperature. When equilibrium is achieved, the partial pressure of A is 1.70 bar. What is the **equilibrium constant** at this temperature?

- A) 1.06  
**B) 1.86**  
 C) 1.21  
 D) 1.81

	$\text{A}_2(\text{g})$	$\rightleftharpoons$	$2 \text{A}(\text{g})$
I	1.50		1.80
C	+x		-2x
E	1.5+x		1.8-2x = 1.7

Since  $1.8-2x = 1.7$ ,  $x = 0.05$ , therefore  $p_{\text{A}_2}$  at equilibrium = 1.55 bar

$$K = p_{\text{A}}^2/p_{\text{A}_2} = (1.7)^2/1.55 = 1.86$$

19. If 1 mole of each the following salts were placed in separate beakers each containing 1 L of water, which one would have the **greatest conductivity**?

- A)  $\text{K}_3\text{PO}_4 \rightarrow 4$  moles of dissolved ions  
**B)  $\text{Fe}_2(\text{SO}_4)_3 \rightarrow 5$  moles of dissolved ions**  
C)  $\text{NH}_4\text{NO}_3 \rightarrow 2$  moles of dissolved ions  
D)  $\text{MgCO}_3 \rightarrow 0$  moles of dissolved ions (insoluble)

- $\text{Fe}_2(\text{SO}_4)_3$  is soluble and generates a higher concentration of ions in solution, resulting in the greatest conductivity.

20. **How many** of the following ionic compounds would generate a **basic solution** when dissolved in water?

$\text{NH}_4\text{Cl}$ ,  $\text{NaHCOO}$ ,  $\text{NaCl}$ ,  $\text{KNO}_3$

- A) 1**  
B) 3  
C) 4  
D) 2

- Consider how these salts would be prepared in acid base reactions that produce water as the only byproduct (or produce no byproducts):



21. The pH of a 1 M aqueous solution of an acid at room temperature is found to be 5.3. What is the relationship between  $K_a$  of the acid and  $K_b$  of its conjugate base?

- A) The relationship between  $K_a$  and  $K_b$  cannot be determined from the given information  
 B)  $K_a = K_b$   
 C)  $K_a < K_b$   
 D)  $K_a > K_b$

	HA	+	H <sub>2</sub> O	$\rightleftharpoons$	H <sub>3</sub> O <sup>+</sup>	+	A <sup>-</sup>
I	1.0		---		0		0
C	-x		---		+x		+x
E	1-x		---		x		x

If the pH is 5.3,  $[H_3O^+] = x = 10^{-5.3} = 5.01 \times 10^{-6}$

$K_a = x^2/(1-x)$ , and since x is very small, this can simplify to  $K_a = x^2 = (5.01 \times 10^{-6})^2 = 2.51 \times 10^{-11}$

$K_a \times K_b = K_w$ , so  $K_b = K_w/K_a = (1 \times 10^{-14})/(2.51 \times 10^{-11}) = 3.98 \times 10^{-4}$

Therefore,  $K_b$  for A<sup>-</sup> is  $> K_a$  for HA (effectively, the reason that HA is only a weak acid is that the conjugate base, A<sup>-</sup> is actually a reasonably good base)

22. Which term best describes BCl<sub>3</sub>?

- A) **Lewis Acid**  
 B) Bronsted-Lowry Acid  
 C) Bronsted-Lowry Base  
 D) Lewis Base

BCl<sub>3</sub> is a Lewis acid.

Boron has just 6 valence electrons, and would much prefer to have 8 (which can be achieved by reaction with a Lewis base).

23. Butanoic Acid (MW = 88.11 g mol<sup>-1</sup>) is a monoprotic acid with  $K_a = 1.5 \times 10^{-5}$ . Calculate the **percent ionization** of butanoic acid when 5.3 g is added to 450 mL of water.

- A) **1.1%**  
 B) 13%  
 C) 0.0020%  
 D) 0.14%

Moles of butanoic acid = 5.3 g / 99.11 g mol<sup>-1</sup> =  $5.35 \times 10^{-2}$  moles

Concentration of butanoic acid solution =  $5.35 \times 10^{-2}$  moles / 0.450 L = 0.1188 M

	HA	+	H <sub>2</sub> O	$\rightleftharpoons$	H <sub>3</sub> O <sup>+</sup>	+	A <sup>-</sup>
I	0.1188		---		0		0
C	-x		---		+x		+x
E	0.1188-x		---		x		x

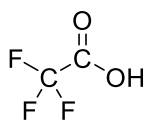
$K_a = 1.5 \times 10^{-5} = x^2 / (0.1188 - x)$ , and we can use the small x approximation, so this becomes:

$$K_a = 1.5 \times 10^{-5} = x^2 / 0.1188$$

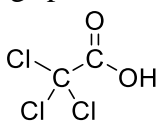
$$x = (1.78 \times 10^{-6})^{1/2} = 1.34 \times 10^{-3}$$

$$\% \text{ ionization} = [\text{H}_3\text{O}^+] / [\text{HA}_{\text{initial}}] = [1.34 \times 10^{-3} / 0.1188] \times 100 \% = 1.12 \%$$

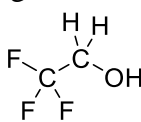
24. Rank the following species according to **increasing**  $K_a$ , from lowest  $K_a$  to highest  $K_a$ .



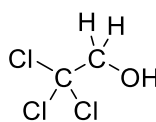
I



II



III



IV

- A) IV < II < III < I  
 B) III < IV < I < II  
 C) I < II < III < IV  
 D) **IV < III < II < I**

Carboxylic acids are more acidic than alcohols (due to resonance stabilization of the conjugate anion).

F<sub>3</sub>CCO<sub>2</sub>H is more acidic than Cl<sub>3</sub>CCO<sub>2</sub>H because the electronegative fluorine atoms make the conjugate base, F<sub>3</sub>CO<sub>2</sub><sup>-</sup>, less basic than Cl<sub>3</sub>CCO<sub>2</sub><sup>-</sup>. For the same reasons, F<sub>3</sub>CCH<sub>2</sub>OH is more acidic than Cl<sub>3</sub>CCH<sub>2</sub>OH.

25. How many **grams of calcium hydroxide** would be required to completely neutralize an 80.0 mL solution of 0.650 M of hydrochloric acid?

- A) **1.93 g**  
B) 0.0260 g  
C) 0.0520 g  
D) 3.86 g



$$\text{Moles of HCl} = 0.08 \text{ L} \times 0.650 \text{ moles L}^{-1} = 0.052 \text{ moles}$$

$$\text{This will react with } (0.052/2) \text{ moles of Ca(OH)}_2 = 0.026 \text{ moles}$$

$$\text{Molecular weight of Ca(OH)}_2 = 74.09 \text{ g mol}^{-1}$$

$$\text{Mass of Ca(OH)}_2 = 0.026 \text{ moles} \times 74.09 \text{ g mol}^{-1} = 1.93 \text{ g}$$

26. **Rank** the following oxyacids from **least acidic** to **most acidic**.



- A)  $\text{HClO} < \text{HBrO} < \text{HClO}_2$   
B)  $\text{HClO}_2 < \text{HClO} < \text{HBrO}$   
C)  $\text{HClO}_2 < \text{HBrO} < \text{HClO}$   
D)  **$\text{HBrO} < \text{HClO} < \text{HClO}_2$**

$\text{HClO} > \text{HBrO}$ , because the more electronegative Cl atom makes  $\text{ClO}^-$  anion less basic than  $\text{BrO}^-$

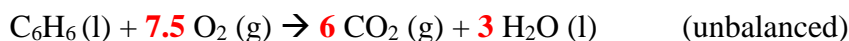
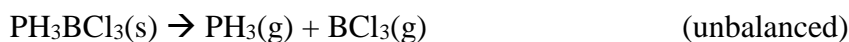
$\text{HClO}_2 > \text{HClO}$ , because there are more resonance structures for  $\text{ClO}_2^-$  relative to  $\text{ClO}^-$ .

27. A gas undergoes a process in which it does 89 J of work on the surroundings and absorbs 567 J of heat. **What is the total change in internal energy, in J, for the gas?**

- A) + 478 J  
B) + 656 J  
C) - 478 J  
D) - 656 J

$$\Delta U = q + w = (+567 \text{ J}) + (-89 \text{ J}) = +478 \text{ J}$$

28. **Balance** the following reactions, and then determine **how many** of these reactions demonstrate work being done *on* the system *by* the surroundings.



- A) 0  
B) **2**  
C) 3  
D) 1

Work is done on a system by the surroundings if  $\Delta n_{\text{gas}}$  is negative

Rxn 1:	1.5 moles of gas	→	1 mole of gas	(negative $\Delta n_{\text{gas}}$ )
Rxn 2:	0 moles of gas	→	2 moles of gas	(positive $\Delta n_{\text{gas}}$ )
Rxn 3:	7.5 moles of gas	→	6 moles of gas	(negative $\Delta n_{\text{gas}}$ )

29. During experiment 2, Cycles of Copper, a student obtains a percent yield/recovery that is more than 100%. Which of the following observations is **NOT a plausible cause of inflated yield** for this experiment?

- A) The final product was slightly damp and smelled of acetone when its mass was recorded.
- B) The student skipped the step that asked them to add HCl to the beaker.
- C) After adding the Zn, the solution was still faintly blue before the copper product was rinsed and dried.**
- D) The actual mass of Cu(s) reacted was 0.2031 g, but the student accidentally used a value of 0.2013 g in their calculations.

- A) A product that is wet with solvent will lead to an inflated yield.
- B) Adding HCl to the beaker removes any remaining solid Zn, and forgetting to do this would lead to a product containing some Zn, giving an inflated yield.
- C) After adding Zn metal, if the solution was still blue, this indicates that not all  $\text{Cu}^{2+}$  was reduced, and this would lead to a **DECREASED** yield.
- D) If the student accidentally used more Cu than they realized at the start of the experiment, this would lead to an inflated yield.

30. During experiment 1, a titration of HCl(aq) with NaOH(aq) was completed. How many of the following mistakes would have **an impact on the accurate determination of the [NaOH]**?

- I. Adding water to the titration flask after pipetting in 10.00 mL of HCl
- II. Allowing some HCl to drip out of the pipette onto the lab bench before transferring to the Erlenmeyer flask
- III. Leaving water in the beaker when collecting the stock HCl solution.

- A) 1
- B) 2**
- C) 3
- D) 0

- I) Adding water to the titration flask after adding the HCl would not change the amount of HCl in the flask, so it would have no effect.
- II) Allowing some HCl to drop out of the pipette would result in less HCl in the Erlenmeyer, so less NaOH will be required to neutralize the solution, and we will think that the concentration of our NaOH solution is higher than it actually is.
- III) Leaving water in the beaker when collecting the stock solution would dilute your sample of the stock solution, again resulting in less HCl in the Erlenmeyer (as above).

- Some general data are provided on this page.
- A Periodic Table with atomic weights is provided on the next page.

$$\text{STP} = 273.15 \text{ K}, 1 \text{ atm}$$

$$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$$

$$h = 6.6256 \times 10^{-34} \text{ Js}$$

$$\text{density}(\text{H}_2\text{O}, \text{l}) = 1.00 \text{ g/mL}$$

$$\text{Specific heat of water} = 4.184 \text{ J / g} \cdot ^\circ\text{C}$$

$$R = 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} = 0.083145 \text{ L bar K}^{-1} \text{ mol}^{-1}$$

$$F = 96485 \text{ C/mol}$$

$$c = 2.9979 \times 10^8 \text{ m/s}$$

$$m_e = 9.109 \times 10^{-31} \text{ kg}$$

$$\Delta H^\circ_{\text{vap}}[\text{H}_2\text{O}] = 44.0 \text{ kJ mol}^{-1}$$

$$1 \text{ bar} = 100.00 \text{ kPa} = 750.06 \text{ mm Hg} = 0.98692 \text{ atm}$$

$$0^\circ\text{C} = 273.15 \text{ K}$$

$$1 \text{ J} = 1 \text{ kg m}^2 \text{ s}^{-2} = 1 \text{ kPa L} = 1 \text{ Pa m}^3$$

$$1 \text{ m} = 10^6 \mu\text{m} = 10^9 \text{ nm} = 10^{10} \text{ \AA}$$

$$1 \text{ cm}^3 = 1 \text{ mL}$$

$$1 \text{ g} = 10^3 \text{ mg}$$

$$1 \text{ Hz} = 1 \text{ cycle/s}$$

De Broglie wavelength:

Hydrogen atom energy levels:

$$\lambda = h / mu = h / p$$

$$E_n = -R_H / n^2 = -2.179 \times 10^{-18} \text{ J} / n^2$$

$$KE = \frac{1}{2}mu^2$$

Nernst Equation:

$$E = E^\circ - \frac{RT}{zF} \ln Q = E^\circ - \frac{0.0257 \text{ V}}{z} \ln Q = E^\circ - \frac{0.0592 \text{ V}}{z} \log_{10} Q$$

Entropy change:  $\Delta S = \frac{q_{\text{rev}}}{T}$

### Solubility Guidelines for Common Ionic Solids

Follow the lower-numbered guideline when two guidelines are in conflict. This leads to the correct prediction in most cases.

1. Salts of group 1 cations and the  $\text{NH}_4^+$  cation are soluble. Except  $\text{LiF}$  and  $\text{Li}_2\text{CO}_3$  which are insoluble.
2. Nitrates, acetates, bicarbonates, and perchlorates are soluble.
3. Salts of silver, lead and mercury (I) are insoluble. Except  $\text{AgF}$  which is soluble.
4. Fluorides, chlorides, bromides, and iodides are soluble. Except Group 2 fluorides which are insoluble
5. Carbonates, phosphates, chromates, sulfides, oxides, and hydroxides are insoluble. Except Group 2 sulfides and hydroxides of  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ , and  $\text{Ba}^{2+}$  which are soluble.).
6. Sulfates are soluble except for those of calcium, strontium, and barium.



Name: \_\_\_\_\_

Student number: \_\_\_\_\_

PERIODIC TABLE OF THE ELEMENTS															
ALDRICH®															
Transition Metals															
I	II	III	IV	V	VI	VII	VIII	IX	X	XI	XII	III	IV	V	VI
1 H 1.0079	2 He 4.0026	3 Li 6.941	4 Be 9.0122	5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180	11 Na 22.990	12 Mg 24.305	13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.066
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.922	34 Se 78.96
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc [98]	44 Ru 101.07	45 Rh 102.91	46 Pd 105.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04
87 Fr [223]	88 Ra 226.03	89 **Ac 227.03	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]
* Lanthanides															
** Actinides															
Atomic weights are based on <sup>12</sup> C = 12 and conform to the 1987 IUPAC report values rounded to 5 significant digits. Numbers in [ ] indicate the most stable isotope.															
[222]															
[210]															
[209]															
[222]															