Name: _____ Student No.:_____

CHMB31H3 – Introduction to Inorganic Chemistry
Term test 2

November 8, 2021

Answer all questions in full.

Value of each question is indicated. Total is ~~64~~ 62 marks.

Aids Allowed*:** You can use lecture notes, lab manual and textbook as only aids. Required equations, graphs and constants you can find in lecture slides and/or textbook**.** ***Any other resources are not allowed

Please, read each question carefully and make sure you have indeed answered every part of each question!!!

Duration: 90 min (1.5 h) + 45 min for test upload

GOOD LUCK!!

Student Name: _____

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1. (~~14~~ *marks*) The following questions are looking into magnesium fluoride, MgF_2.
   a. Using Hess cycle calculate ΔH_{latt} for MgF_2. Besides the thermodynamic data you can find on lecture slides, you'll also need the following: $\Delta_f H(MgF_2)$ = −1124.2 kJ mol^{-1}, $BDE(F_2)$ = 158.8 kJ mol^{-1} and $EA(F(g))$ = −328 kJ mol^{-1}.
   b. When we had an overview of Group 2 halides, we have seen that MgF_2 has TiO_2-type lattice. Without performing any calculations, decide which one, MgF_2 or TiO_2, has a more stable lattice (or lower lattice energy). Explain your reasoning.

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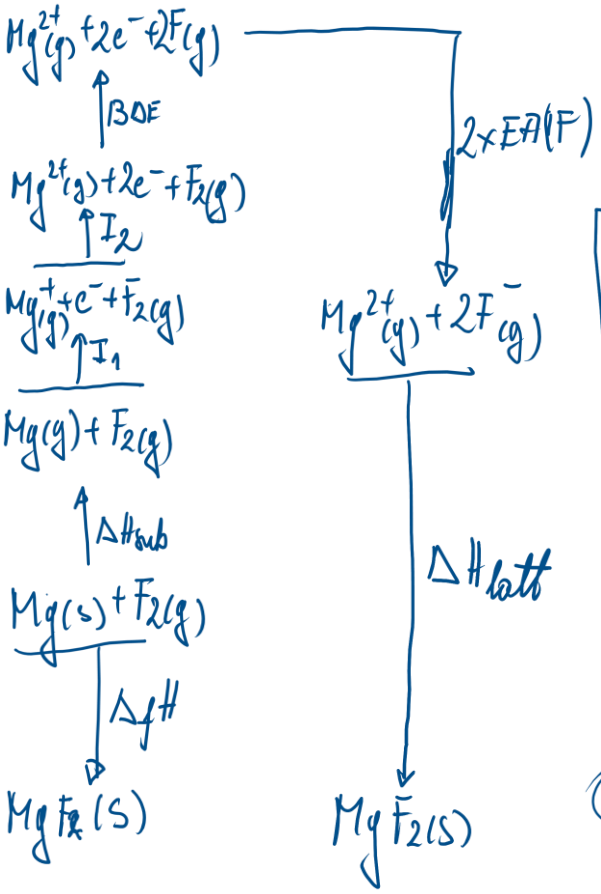
Answer:

a.



Mg(s) ⟶ Mg(g); ΔH_{sub} = +146 kJ mol^{-1} ✓

Mg(g) ⟶ $Mg^+(g) + e^-$; I_1 = +737.7 kJ mol^{-1} ✓

$Mg^+(g) \longrightarrow Mg^{2+}(g) + e^-$; I_2 = +1451 kJ mol^{-1} ✓

$F_2(g) \longrightarrow 2F(g)$; $BDE(F_2)$ = +158.8 kJ mol^{-1} ✓

$2F(g) + 2e^- \longrightarrow 2F^-(g)$; 2EA = 2×(-328)kJ mol^{-1} ✓

$Mg(g) + 2F^-(g) \longrightarrow MgF_2(s)$; ΔH_{latt} = ? ✓

$Mg(s) + F_2(g) \rightarrow MgF_2(s)$, $\Delta_f H$ = -1124.2 kJ mol^{-1} ✓

→ each of these is 1 mark (the same for the scheme on the left: each step is 1 mark) ×7

$\Delta_f H = \Delta H_{sub} + I_1 + I_2 + BDE - 2\times EA - \Delta H_{latt}$ (1 mark)

$\Delta H_{latt} = 2961.7$ kJ/mol (1 mark) ± 1 mark units

b. TiO_2 (1 mark), because the charges are higher (1 mark)

2. (*17 marks*) Consider H_2S, SiH_4 and PH_3.
   a. Classify these three hydrides.
   b. For each hydride calculate partial charge, δ, for each elements using the following electronegativity values: H, 2.2; S, 2.58; Si, 1.9 and P, 2.19.
   c. Based on your results for part b., decide which of theses hydrides is the easiest to deprotonate. Explain your reasoning.

Answer:

a.

H_2S: molecular, electron rich

SiH_4: molecular, electron precise

PH_3: molecular, electron rich

b. $\delta_{LL} = \text{Number of valence electrons on X} - \text{Number of lone electrons on X} - 2 \times \sum_{bonds}\left(\frac{\chi_x}{\chi_x+\chi_y}\right)$

H_2S:

$$\delta_{LL}(S) = 6 - 4 - 2 \times 2 \times \frac{2.58}{2.58 + 2.2} = -0.16$$

$$\delta_{LL}(H) = 1 - 0 - 2 \times 1 \times \frac{2.2}{2.58 + 2.2} = +0.06$$

(Alternatively, H can be calculated as $-\delta(S)/2 = +0.06$)

SiH_4:

$$\delta_{LL}(Si) = 4 - 0 - 2 \times 4 \times \frac{1.9}{1.9 + 2.2} = +0.29$$

$$\delta_{LL}(H) = 1 - 0 - 2 \times 1 \times \frac{2.2}{1.9 + 2.2} = -0.07$$

(Alternatively, H can be calculated as $-\delta(Si)/4 = -0.07$

PH_3:

$$\delta_{LL}(P) = 5 - 2 - 2 \times 3 \times \frac{2.19}{2.19 + 2.2} = +0.01 \approx 0$$

$$\delta_{LL}(H) = 1 - 0 - 2 \times 1 \times \frac{2.2}{2.19 + 2.2} = -0.003 \approx 0$$

(Alternatively, H can be calculated as –δ(P)/3 = –0.003 (0 is also acceptable)

c.

It is H_2S since it contains the most positive H atom

3. (*6 marks*) Let's have a look at the hydrides of group 2.
   a. Looking at the structures of BeH_2 and MgH_2 covered in the lectures, determine the coordination number for Be^{2+} and Mg^{2+} in the hydride structures.
   b. We said that the structures of other Group 2 hydrides are more complex. In these structures the coordination number of cations is nine. Why would these cations have such a high coordination number in hydride structures?
   c. We said that generally the reactivity of Group 2 hydrides is lower than the reactivity of Group 1 hydrides. But how is the reactivity of Group 2 hydrides (for example with H_2O) changing down the group 2? Explain your reasoning.

Answer:

a. For Be it is 2 and for Mg it is 6

b. Cations are larger and can have greater coordination number

c. The reactivity would increase because the hydrides are more ionic (or because reactivites increase down the group)

4. (3 marks) Although beryllate anion, $[Be(OH)_4]^{2-}$ has been known for a long time, its tetrahedral structure was experimentally confirmed only relatively recently – in 2011. The compound used for the experiment was $Sr[Be(OH)_4]$. Explain how you would prepare this compound from $BeCl_2$ and write a balanced equation(s) for its synthesis.

Answer:

This can be answered in several ways:

1. Stepwise:

$BeCl_2 + Sr(OH)_2 = Be(OH)_2 + SrCl_2$ (it is also acceptable to use NaOH here to make $Be(OH)_2$)

$Be(OH)_2 + Sr(OH)_2 = Sr[Be(OH)_4]$

2. One reaction:

$BeCl_2 + 2Sr(OH)_2 = Sr[Be(OH)_4 + SrCl_2$

5. (9 *marks*) During our excursion to group 1 chemistry we mentioned alkalides.
   a. Using the ionization potentials given in the lecture and electron affinities given below, calculate ΔH for the following reaction (M is group 1 metal):

$$2M(g) \rightarrow M^+(g) + M^-(g)$$

Electron affinities for Group 1 metals are (all values are in kJ mol^{-1}): Li, −59.6; Na, −52.9; K, −48.4; Rb −46.9 and Cs, −45.5.

   b. Based on your results, determine which Group 1 element is most likely to produce alkalide following the above reaction. Explain.
   c. Using the same thermodynamic data, decide which compound is more likely to form Cs^+Na^- or Na^+Cs^-? Which other thermodynamic value is important to consider here?

Answer:

a. This is all the same procedure:

$\Delta H = I_1 + EA$ for all of them and they have to find I_1 in the lecture slides:

Li: +520.2 kJ mol^{-1} + (–59.6 kJ mol^{-1}) = 460.6 kJ mol^{-1}

Na: +495.8 kJ mol^{-1} + (–52.9 kJ mol^{-1}) = 442.9 kJ mol^{-1}

K: +418.8 kJ mol^{-1} + (–48.4 kJ mol^{-1}) = 370.4 kJ mol^{-1}

Rb: +403 kJ mol^{-1} + (–46.9 kJ mol^{-1}) = 356.1 kJ mol^{-1}

Cs: +375.7 kJ mol^{-1} + (–45.5 kJ mol^{-1}) = 330.2 kJ mol^{-1}

b. Cs, since ΔH is the lowest.

c. For Cs^+Na^- : $I_1(Cs) + EA(Na) = +375.7\ kJ\ mol^{-1} + (-52.9\ kJ\ mol^{-1}) = 322.8\ kJ\ mol^{-1}$

For Na^+Cs^- : $I_1(Na) + EA(Cs) = +495.8\ kJ\ mol^{-1} + (-45.5\ kJ\ mol^{-1}) = 450.6\ kJ\ mol^{-1}$

The first compound is more likely to form.

We would need ΔH_{latt} as well

6. (4 *marks*) The first ionization potential for Na is given in the Group 1 lecture. The second ionization potential for Na is 4562 kJ mol^{-1}. The I_1 and I_2 for Mg are given in lecture on Group 2 metals. Explain:
    a. Why is I_1 for Na lower than I_1 for Mg?
    b. Why is I_2 for Na higher than I_2 form Mg?

Answer:

a. The Z_{eff} is increasing from Na to Mg making I_1 for Mg higher

b. The I_2 for Na removes core electrons while I_2 for Mg removes second valence electron

7. (8 *marks*) Answer the following questions:
    a. Which cryptand would you expect to form the most stable complex with Ca^{2+}? You have to be specific – look at cryptands in lecture on group 1 chemistry to pick one specific crypt compound. Explain your choice.
    b. Why is barium forming a peroxide, not an oxide, when reacted with O_2?
    c. What is the product of reaction between BaO_2 and water?

d. What would be the product when radium, Ra, at the bottom of group 2, reacts with O_2? Explain your reasoning.

Answer:

a. Ca^{2+} has a similar ionic radius as Na^{+} so I would expect 2.2.1-crypt to be the best

b. Ba^{2+} is a large cation and needs a large O_2^{2-} to stabilize the product (or lattice)

c. It is $Ba(OH)_2$ and H_2O_2

d. It would likely be superoxide (peroxide is also acceptable) because Ra^{2+} is the largest cation of Group 2

8. (6 marks) Determine the product and write balanced chemical equations for the following reactions
   a. Magnesium hydride and hydrochloric acid
   b. Strontium carbide and ethanol
   c. Calcium nitride and water

Answer:

a. $MgH_2 + 2HCl = MgCl_2 + 2\ H_2$

b. $SrC_2 + 2CH_3CH_2OH = Sr(OCH_2CH_3)_2 + H_2C_2$

c. $Ca_3N_2 + 6H_2O = 3Ca(OH)_2 + 2NH_3$

- END OF TEST –