

Name: _____

Student number: _____

Chemistry 1E03

Test 2 Solutions

Nov. 10, 2017

McMaster University

VERSION 1

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Duration: 120 minutes

1. Which of the following statements are **FALSE** according to the shapes predicted by the VSEPR model? (Central atoms are underlined.)

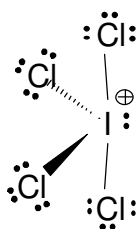
- (i) AsH₃ and BrO₃⁻ have the same shape.
- (ii) ICl₄⁺ is tetrahedral.
- (iii) ICl₄⁻ is square-planar.
- (iv) OF₂ has a smaller permanent dipole moment than XeF₂.
- (v) NO₂ is a non-polar molecule.

- A) ii, iii
- B) i, iii, iv
- C) i, ii
- D) ii, iv, v**
- E) ii, v

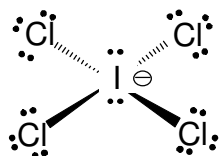
(i) AsH₃ and BrO₃⁻ are both in VSEPR class, AX₃E, and are both trigonal pyramidal:



(ii) ICl₄⁺ is in VSEPR class, AX₄E, and is seesaw shaped:

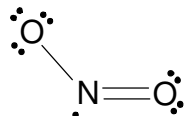


(iii) ICl₄⁻ is in VSEPR class, AX₄E₂, and is square planar:



(iv) OF_2 is AX_2E_2 – i.e., bent shaped – whereas XeF_2 is AX_2E_3 which is linear. The bond dipoles cancel in XeF_2 . It has zero molecular dipole moment, which is smaller than that of OF_2 .

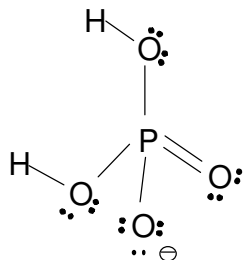
(v) NO_2 is AX_2E . It is bent, and therefore polar.



2. How many **non-bonding electrons** are there in the charge-minimized Lewis structure for the dihydrogen phosphate ion, H_2PO_4^- ? (In the Lewis structure P is the central atom, and is bonded only to O).

- A) 6
B) 10
C) 18
D) 20
E) 14

There are 9 non-bonding electron pairs – i.e., 18 non-bonding electrons.



3. According to the VSEPR model, which one of the following molecules should be **linear**? (Central atoms are underlined)

- A) $\text{H}_2\text{C}\underline{\text{O}}$
B) $\underline{\text{B}}\text{F}_3$
C) $\underline{\text{S}}\text{O}_2$
D) $\text{H}\underline{\text{C}}\text{N}$
E) $\text{H}_2\underline{\text{S}}$

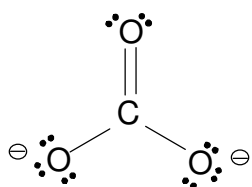
- A) AX_3 - trigonal planar
B) AX_3 - trigonal planar
C) AX_2E – bent
D) AX_2 - linear

E) AX₂E – bent

4. **Rank** the species CO₂, CO, CO₃²⁻ and HCO₂⁻ in order of **increasing bond order** of the carbon-oxygen bonds. For each species, carbon is the central atom.

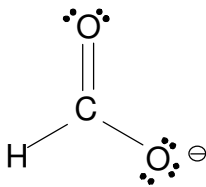
- A) CO < CO₂ < HCO₂⁻ < CO₃²⁻
 B) HCO₂⁻ < CO₃²⁻ < CO₂ < CO
 C) CO < CO₃²⁻ < HCO₂⁻ < CO₂
D) CO₃²⁻ < HCO₂⁻ < CO₂ < CO
 E) CO₃²⁻ < HCO₂⁻ < CO < CO₂

CO has a triple bond – i.e., bond order equals 3. CO₂ has two double bonds – i.e., bond order equals 2. HCO₂⁻ has two equivalent CO bonds with average bond order equal to 3/2. CO₃²⁻ has three equivalent CO bonds with average bond order equal to 4/3.



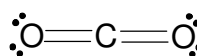
three resonance structures

average bond order = 4/3



two resonance structures

average bond order = 3/2



5. Which one of the following molecules has **no** net dipole moment? (Central atoms are underlined)

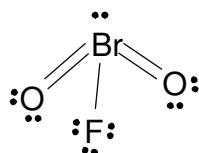
- A) NF₃
B) SO₃
 C) NO₂
 D) CH₃Cl
 E) H₂Se

- A) AX₃E - trigonal pyramidal – net dipole moment
 B) AX₃ - trigonal planar – no net dipole
 C) AX₂E – bent – net dipole moment
 D) AX₄ - tetrahedral. However, the CCl bond dipole is different from the CH bond dipoles – net dipole moment
 E) AX₂E – bent – net dipole moment

6. Which of the following statements are **FALSE** about the BrO_2F molecule? (Br is the central atom)

- (i) It has a permanent dipole moment.
- (ii)** It is T-shaped about Br.
- (iii) There is one lone pair of electrons on Br.
- (iv) The octet rule is violated at Br.
- (v) The oxidation number of Br is +5.
- (vi)** The average Br-O bond order is 1.5.

- A) iv, vi
B) ii, vi
C) iii, iv
D) ii, iv
E) i, v



7. During experiment 2, Cycles of Copper, a student obtains a percent yield/recovery that is less than 100%. Which of the following observations is **NOT a plausible explanation** for this result?

- A)** The final product was slightly damp and smelled of acetone when its mass was recorded.
- B) Small amounts of CuO(s) were lost during the decanting step.
- 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 0.2013 g, but the student accidentally used a value of 0.2031 g in their calculations.
- E) Despite the addition of the H_2SO_4 , some black precipitate was present when Zn(s) was added to the reaction beaker.
- A) This error would give too large a mass of final product – NOT a possible explanation for < 100% recovery.
- B) Losing Cu (in the form of CuO) is a possible explanation for < 100% recovery.
- C) If not all Cu^{2+} is reduced, there will be less recovery at the end of cycle.
- D) If the initial mass of copper is recorded as higher than what is actually used, the result will be a calculated recovery rate < 100%.
- E) The black color indicates that not all the CuO(s) is dissolved and reacted with sulfuric acid. This means there is less Cu^{2+} available to react with Zn, and final recovery is less than 100%.

8. When the following pairs of solutions are mixed, which produces the **strongest electrolyte** solution?

- A) $\text{Zn}(\text{NO}_3)_2$ (1 M) + AgClO_4 (2 M)
 B) $\text{Ba}(\text{OH})_2$ (3 M) + CuSO_4 (3 M)
 C) $\text{Pb}(\text{CH}_3\text{COO})_2$ (1 M) + NaI (2 M)
 D) KCl (1 M) + NH_4NO_3 (1 M)
 E) HCl (2 M) + NaOH (2 M)

First, determine the nature of the final solution. The strongest electrolyte solution has the highest concentration of ions. To simplify the discussion, suppose there is 1 L of each solution.

A) $\text{Zn}(\text{NO}_3)_2$ (1 M) contains 1 mol Zn^{2+} and 2 mol NO_3^- . AgClO_4 (2 M) has 4 mol of ions. There is no reaction when the solutions are added. Consequently, there are 7 mol of ions in the resulting 2 L solution.

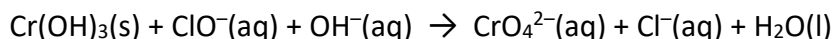
B) BaSO_4 and $\text{Cu}(\text{OH})_2$ are both insoluble (see solubility rules in info sheet). Since the ions appear in stoichiometric proportions, no ions remain in solution.

C) PbI_2 is insoluble. This leaves 2 mol CH_3COO^- and 2 mol Na^+ in 2 L solution.

D) 4 mol ions in 2 L solution.

E) The H^+ and OH^- neutralize, leaving 2 mol Cl^- and 2 mol Na^+ in 2 L solution.

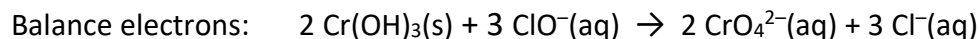
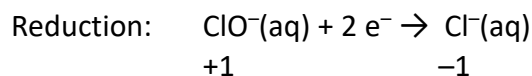
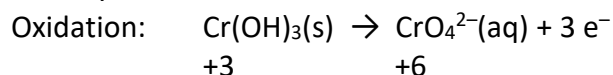
9. Balance the following redox reaction in **basic** solution.



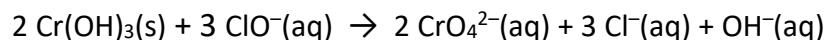
When this has been done correctly, the stoichiometric **coefficients** for the **reactant** species, in order from **left** to **right** are as follows:

- A) 1, 2, 3
 B) 2, 4, 6
 C) 3, 4, 5
 D) 2, 3, 4
 E) 2, 3, 7

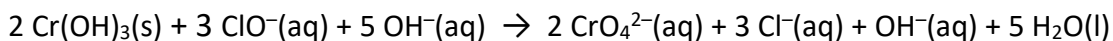
Identify the half reactions.



Balance O's (in base):

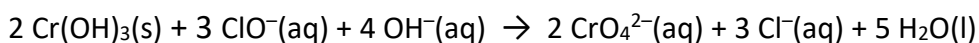
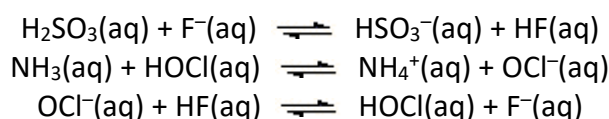


Balance H's (in base):



Name: _____

Student number: _____

Cancel redundant OH^- :10. The following equilibria are **all** shifted towards products:Order the **acids** in these reactions according to **increasing** strength as acids.

- A) $\text{H}_2\text{SO}_3 < \text{HOCl} < \text{NH}_4^+ < \text{HF}$
B) $\text{HSO}_3^- < \text{F}^- < \text{OCl}^- < \text{NH}_3$
C) $\text{OCl}^- < \text{NH}_4^+ < \text{HF} < \text{HSO}_3^-$
D) $\text{HOCl} < \text{NH}_4^+ < \text{H}_2\text{SO}_3 < \text{HF}$
E) $\text{NH}_4^+ < \text{HOCl} < \text{HF} < \text{H}_2\text{SO}_3$

Acid-base equilibria are always shifted away from the stronger acid. Identify the acid on both sides of reaction. The reactant acid is stronger than the product acid. The first reaction tells us that $\text{H}_2\text{SO}_3 > \text{HF}$. The second reaction tells us that $\text{HOCl} > \text{NH}_4^+$. The last reaction tells us that $\text{HF} > \text{HOCl}$.

11. Identify the **TRUE** statement about the reactions of metallic Zn:

- A) Zn(s) dissolves in water at room temperature.
B) Zn(s) dissolves in hydrochloric acid and chlorine gas is produced.
C) Zn(s) dissolves in nitric acid and hydrogen gas is produced.
D) Zn(s) dissolves in nitric acid and nitric oxide is produced.
E) In the reaction with copper sulphate solution, Zn(s) acts as the oxidizing agent.
- A) Metallic zinc does not dissolve in water.
B) Metallic zinc reacts with hydrochloric acid, dissolving and producing a zinc chloride solution. Hydrogen gas is produced. Chlorine gas is NOT produced.
C) and D) Metallic zinc reacts with nitric acid, dissolving and producing a zinc nitrate solution. Hydrogen gas is produced.
E) In the reaction with copper sulphate solution, Zn(s) acts as the reducing agent – it gets oxidized to $\text{Zn}^{2+}(\text{aq})$.

12. Which of the following is a **Brønsted-Lowry acid-base** reaction?

- A) $\text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
B) $\text{Ca}(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
C) $\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
D) $\text{Cu}(\text{s}) + 4\text{HNO}_3(\text{aq}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{aq}) + 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
E) $2\text{K}_3\text{PO}_4(\text{aq}) + 3\text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{Pb}_3(\text{PO}_4)_2(\text{s}) + 6\text{KNO}_3(\text{aq})$

A Brønsted-Lowry acid-base reaction is an H^+ transfer. Look for something that lost an H^+ and something that gained it. B and D are redox reactions. E is a precipitation. C is a decomposition.

13. Which **one** of the following will oxidize $\text{H}_2\text{S}(\text{g})$?

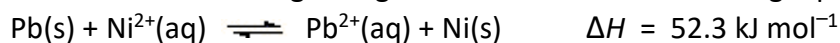
- A) $\text{Na}(\text{s})$
B) $\text{HBr}(\text{aq})$
C) $\text{NH}_3(\text{g})$
D) $\text{KMnO}_4(\text{aq})$
E) $\text{H}_2(\text{g})$

An oxidizing agent has an atom in a higher oxidation state, which can go to a lower oxidation state.

14. For the gas phase reaction, $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$, the expression for the **reaction quotient**, Q , is as follows:

- A) $Q = P(\text{NO}_2) / P(\text{NO}) P(\text{O}_2)$
B) $Q = P(\text{NO}_2)^2 / P(\text{NO})^2 P(\text{O}_2)$
C) $Q = P(\text{NO})^2 P(\text{O}_2) / P(\text{NO}_2)^2$
D) $Q = P(\text{NO}_2)^2$
E) $Q = P(\text{NO})^2 P(\text{O}_2)$

15. Which of the following changes will **not** affect the following equilibrium?



- i. Decrease the temperature
- ii. Add $\text{Pb}(\text{NO}_3)_2(\text{aq})$
- iii. Add Ni(s)
- iv. Add $\text{NaI}(\text{aq})$
- v. Add $\text{H}_2\text{O(l)}$

- A) iii and iv
- B) iii and v**
- C) i and ii
- D) i and iv
- E) ii and v

- i. Decreasing temperature shifts equilibrium to reactants (the exothermic direction).
- ii. Adding $\text{Pb}(\text{NO}_3)_2(\text{aq})$ shifts equilibrium to reactants – a product is added.
- iii.** Adding Ni(s) has no effect. Solids have fixed activity of 1.
- iv. Adding $\text{NaI}(\text{aq})$ causes $\text{PbI}_2(\text{s})$ to precipitate, shifting equilibrium to products – a product is removed.
- v.** Adding $\text{H}_2\text{O(l)}$ dilutes all ion concentrations. However, $Q = [\text{Pb}^{2+}]/[\text{Ni}^{2+}]$ is unaffected, as both concentrations are reduced by the same factor. Here, no affect arises because $\Delta n_{\text{aq}} = 0$.

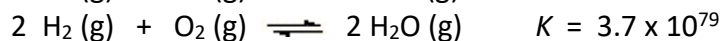
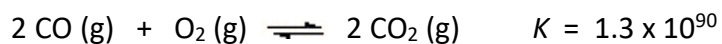
16. Choose the **FALSE** statements regarding chemical equilibrium.

- (i) When the reaction quotient is greater than the equilibrium constant, there is net reverse reaction.
- (ii)** Increasing the temperature always increases the equilibrium constant.
- (iii)** When the system has reached equilibrium, the molecules stop reacting.
- (iv) When the system has reached equilibrium, the concentrations of products and reactants stop changing.

- A) iii, iv
- B) ii, iii
- C) ii, iv
- D) i, ii
- E) i, iii

- ii. Increasing temperature increases the equilibrium constant of endothermic reactions, but decreases the equilibrium constant of exothermic reactions.
- iii. Reaction continues at equilibrium. However, there is no NET reaction as the rate of forward reaction equals the rate of reverse reaction.

17. At 25°C, the following reactions have the equilibrium constants noted to the right of their equations:



Determine K for the following reaction:

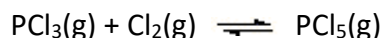


- A) 3.6×10^{10}
B) 1.9×10^5
 C) 3.2×10^5
 D) 5.3×10^{-6}
 E) 1.8×10^{10}

The target reaction equals $\frac{1}{2}$ x first reaction – $\frac{1}{2}$ x second reaction. Therefore,

$$K = K_1^{1/2} K_2^{-1/2} = (1.3 \times 10^{90})^{1/2} (3.7 \times 10^{79})^{-1/2}$$

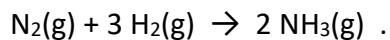
18. At 250°C, a reaction vessel contains $\text{PCl}_3(\text{g})$, $\text{Cl}_2(\text{g})$ and $\text{PCl}_5(\text{g})$ at equilibrium. The associated partial pressures are 0.867 bar, 0.867 bar and 1.34 bar, in the given order. Determine the **equilibrium constant, K** , for the reaction,



- A) 0.855
 B) -1.30
 C) 0.561
 D) 9.71
E) 1.78

$$\begin{aligned}
 K &= P[\text{PCl}_5] / (P[\text{PCl}_3] P[\text{Cl}_2]) \\
 &= 1.34 / (0.867 \times 0.867) = 1.78
 \end{aligned}$$

19. A vessel is filled with 1.00 bar partial pressure of nitrogen gas, and 3.00 bar hydrogen gas, and allowed to equilibrate at 573 K. Ammonia forms according to the balanced reaction,



$K = 0.0042$ for this reaction at 573 K. What is the **partial pressure** of ammonia (in bar) at **equilibrium**?

- A) 0.069
B) 0.26
 C) 0.79
 D) 1.19
 E) 1.7

	$\text{N}_2(\text{g})$	+	$3 \text{H}_2(\text{g})$	\rightarrow	$2 \text{NH}_3(\text{g})$
I	1.00		3.00		0
C	$-x$		$-3x$		$2x$
E	$1.00-x$		$3.00-3x$		$2x$

$$\begin{aligned}
 K &= 0.0042 \\
 &= \frac{(2x)^2}{(1.00-x)(3.00-3x)^3} \\
 &= \frac{4}{3^3} \frac{x^2}{(1.00-x)^4} \\
 &= \left(\frac{2}{3^{3/2}} \frac{x}{(1.00-x)^2} \right)^2
 \end{aligned}$$

Taking the square root of this equation gives

$$0.0042^{1/2} = \frac{2}{3^{3/2}} \frac{x}{(1.00-x)^2}$$

or

$$\begin{aligned}
 (1.00 - 2x + x^2) &= 0.0042^{-1/2} \frac{2}{3^{3/2}} x \\
 &= 5.94x
 \end{aligned}$$

or

$$1.00 - 7.94x + x^2 = 0$$

which has solutions,

$$x = 7.81 \text{ or } 0.128$$

$x = 0.128$ gives the equilibrium partial pressure of ammonia, $P[\text{NH}_3] = 2x = 0.26 \text{ bar}$.

20. The first step in the industrial manufacture of nitric acid is the combustion of ammonia. Use the standard formation enthalpies given to calculate the standard enthalpy change, ΔH° (in kJ mol^{-1}), for the reaction:



Substance	$\text{NH}_3(\text{g})$	$\text{NO}_2(\text{g})$	$\text{H}_2\text{O}(\text{l})$
$\Delta H_f^\circ / \text{kJ mol}^{-1}$	-46	+34	-286

- A) -1764
 B) -1668
 C) -824
 D) -206
 E) -1396

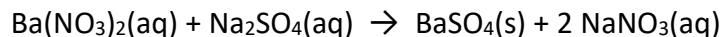
$$\begin{aligned} \Delta H^\circ &= 4 \Delta H_f^\circ[\text{NO}_2(\text{g})] + 6 \Delta H_f^\circ[\text{H}_2\text{O}(\text{l})] - 4 \Delta H_f^\circ[\text{NH}_3(\text{g})] \\ &= 4 \times 34 + 6 \times (-286) - 4(-46) \\ &= -1396 \text{ kJ mol}^{-1} \end{aligned}$$

21. In an ice calorimeter, 15.0 mL of 2.00 M $\text{Ba}(\text{NO}_3)_2(\text{aq})$ are added to 15.0 mL of 2.00 M $\text{Na}_2\text{SO}_4(\text{aq})$. The reaction causes 2.34 grams of ice to melt. What is ΔH° (in kJ mol^{-1}) for the **precipitation of BaSO_4** at 0°C ? $\Delta H_{\text{fus}}^\circ(\text{H}_2\text{O}) = 333 \text{ J g}^{-1}$.

- A) +78.1
 B) +26.0
 C) -78.1
 D) -142
 E) -26.0

$$\begin{aligned} \text{Heat released by reaction} &= \text{mass of ice melted} \times \text{enthalpy of fusion (per g)} \\ q &= 2.34 \text{ g} \times 333 \text{ J g}^{-1} = 779 \text{ J} \end{aligned}$$

Balanced reaction:



$$\text{Extent of reaction} = n = 0.015 \text{ L} \times 2.00 \text{ mol L}^{-1} = 0.030 \text{ mol}$$

Enthalpy of reaction =

$$\begin{aligned} \Delta H^\circ &= -q / n \\ &= -779 \text{ J} / 0.030 \text{ mol} = -26.0 \text{ kJ mol}^{-1} \end{aligned}$$

22. Choose the one **FALSE** statement from among the following:

- A) $\Delta H_f^\circ[\text{CO}_2(\text{g})] > \Delta H_f^\circ[\text{CO}_2(\text{s})]$
- B) ΔH_f° for some compounds is positive, while for others it is negative.
- C**) Doing work on a gas increases its volume.
- D) All diatomic molecules have molar heat capacities that are larger than their specific heat capacities.
- E) The enthalpy change for $\text{O}_2(\text{g}) \rightarrow 2 \text{O}(\text{g})$ is positive.

- A) The gas phase of any substance has more enthalpy than the corresponding solid phase.
- B) ΔH_f° is simply the enthalpy of a substance relative to the enthalpy of the elements in the standard state – it can be positive or negative.
- C) Doing work on a gas requires decreasing its volume.
- D) Molar heat capacity is just molar mass (in g mol^{-1}) \times specific heat capacity (in $\text{J K}^{-1} \text{g}^{-1}$). Since all diatomic molecules have a molar mass greater than 1 (H_2 is the lightest with a molar mass of 2 g mol^{-1}), molar heat capacity is greater than specific heat capacity for all diatomic molecules.
- E) Breaking bonds is endothermic.

23. Given $\Delta H_f^\circ[(\text{NF}_3(\text{g}))] = -132 \text{ kJ mol}^{-1}$ and the bond enthalpy (D) data, $D(\text{N}_2) = 946 \text{ kJ mol}^{-1}$ and $D(\text{F}_2) = 159 \text{ kJ mol}^{-1}$, which of the following statements is/are **FALSE**?

- (i) The average N-F bond enthalpy in NF_3 is 281 kJ mol^{-1} .
- (ii) $\Delta H_f^\circ[(\text{NF}_3(\text{g}))] > \Delta H_f^\circ[(\text{NF}_3(\text{l}))]$
- (iii) $\Delta H_f^\circ[(\text{F}(\text{g}))] = 159 \text{ kJ mol}^{-1}$

- A) ii
- B) i, iii
- C) ii, iii
- D) i
- E) iii

i) The average N-F bond enthalpy, $D(\text{NF})$, in NF_3 is ΔH° for
 $1/3 \text{ NF}_3(\text{g}) \rightarrow 1/3 \text{ N}(\text{g}) + \text{F}(\text{g})$

Consequently, $D(\text{NF}) = \Delta H_f^\circ[\text{F}(\text{g})] + 1/3 \Delta H_f^\circ[\text{N}(\text{g})] - 1/3 \Delta H_f^\circ[(\text{NF}_3(\text{g}))]$

$\Delta H_f^\circ[\text{F}(\text{g})] = 1/2 D(\text{F}_2) = 159/2 \text{ kJ mol}^{-1}$ and $\Delta H_f^\circ[\text{N}(\text{g})] = 1/2 D(\text{N}_2) = 946/2 \text{ kJ mol}^{-1}$

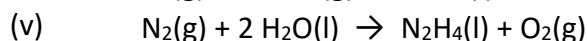
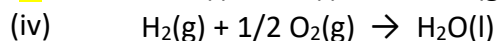
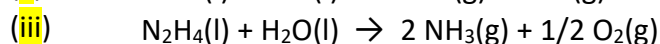
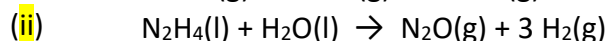
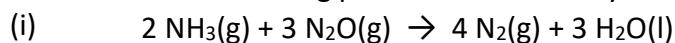
Therefore,

$$\begin{aligned} D(\text{NF}) &= 159/2 + 946/6 - 1/3 (-132) \text{ kJ mol}^{-1} \\ &= 281 \text{ kJ mol}^{-1} \end{aligned}$$

ii) The enthalpy of the gas phase of any substance is greater than the enthalpy of the corresponding liquid phase.

iii) $\Delta H_f^\circ[\text{F}(\text{g})] = 1/2 D(\text{F}_2) = 159/2 \text{ kJ mol}^{-1}$

24. In which of the following processes does the system do work on the surroundings?



- A) all
 B) i, ii, iii, v
 C) i, iv
 D) none
 E) ii, iii

Look for $\Delta n_{\text{gas}} > 0$ (i.e., the system expands, and does work on its surroundings).

25. Fructose, $\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$, is a sugar closely related to glucose. A 0.755 g sample of fructose was combusted with excess oxygen in a bomb calorimeter, containing 500.0 g of water. The heat capacity of the empty calorimeter was 208 J/K. The temperature of the calorimeter and the water rose from 22.00°C to 27.12°C due to the combustion reaction, which formed $\text{CO}_2(\text{g})$ and liquid water. What is the **energy change, ΔU** (in kJ mol^{-1}), for the combustion of one mole of fructose?

- A) -804
 B) -2810
 C) -15600
 D) +254
 E) +520

Heat released by reaction = heat capacity of calorimeter \times temperature change

Heat capacity of calorimeter = $208 \text{ J K}^{-1} + 500.0 \text{ g} \times 4.18 \text{ J K}^{-1}\text{g}^{-1} = 2298 \text{ J K}^{-1}$

$$q = 2298 \text{ J K}^{-1} \times 5.12 \text{ K} = 11.77 \text{ kJ}$$

Extent of reaction = n = number of moles of fructose combusted = $0.755 \text{ g} / 180.16 \text{ g mol}^{-1}$
 = 0.004191 mol

Enthalpy of reaction =

$$\Delta H^\circ = -q / n$$

$$= -2810 \text{ kJ mol}^{-1}$$