1. The solution key will be posted on the web today.
2. You must respond to all exercises.
3. Periodic Table, Electronegativity Table and relevant values/equations are at the end.
4. A few scratch sheets are at the very end. Report in the booklet the minimum amount of calculations to show your reasoning.

Read carefully:

Cellular phones, unauthorized electronic devices or course notes (unless an open-book exam) are not allowed during this exam. Phones and devices must be turned off and put away in your bag. Do not keep them in your possession, such as in your pockets. If caught with such a device or document, the following may occur: academic fraud allegations will be filed which may result in your obtaining a 0 (zero) for the exam.

By signing below, you acknowledge that you have read and ensured that you are complying with the above statement.

Signature: ________________________________
1. (3 points)
Most of the ammonia produced by the Haber process is used as fertilizer. A second important use of NH₃ is in the production of nitric acid, a top-15 industrial chemical. Nitric acid is produced by a three-step synthesis called the Ostwald process:

\[
4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O} \\
2 \text{NO} + \text{O}_2 \rightarrow 2 \text{NO}_2 \\
3 \text{NO}_2 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3 + \text{NO}
\]

The NO is recycled so that every mole of ammonia theoretically yields 1 mole of nitric acid. Starting with \(7.50 \times 10^2\) kg of ammonia, what mass of nitric acid can be produced if the average yield of the three steps is 94.5% efficient?

In a multiple-step synthesis, multiply the yields of all steps to determine the overall yield. In this case, there are three steps, each with a yield of 94.5%, so the overall fractional yield is \((0.945)(0.945)(0.945) = (0.945)^3 = 0.844\). Use this value and the starting amount in moles to determine how many moles of product will form. Do the usual mole–mass conversions:

We are told that every mole of ammonia theoretically produces 1 mole of nitric acid.

\[
\text{Moles of NH}_3 = 7.50 \times 10^2 \text{ kg} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol}}{17.03 \text{ g}} \right) = 4.404 \times 10^4 \text{ mol} = \text{moles HNO}_3
\]

To determine the amount of nitric acid actually produced, multiply the fractional yield by the theoretical yield:

Moles of nitric acid produced = \((4.404 \times 10^4 \text{ mol})(0.844) = 3.716 \times 10^4 \text{ mol}\)

Finally, convert from moles to kilograms to obtain the mass of nitric acid produced:

\[
\text{Mass produced} = 3.716 \times 10^4 \text{ mol} \left( \frac{63.0 \text{ g}}{1 \text{ mol}} \right) \left( \frac{10^{-3} \text{ kg}}{1 \text{ g}} \right) = 2.34 \times 10^3 \text{ kg}
\]

2. (2 points)
Acrylonitrile is an important building block for synthetic fibres and plastics. The compound is synthesized from propene in the following reaction:
How many kilograms of acrylonitrile can be prepared from 1.50 × 10³ kg of propene, 6.80 × 10² kg of ammonia, and 1.92 × 10³ kg of oxygen.

The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the mass of the product formed. Starting amounts are in kilograms, so it will be convenient to work with 10³ mol amounts. The balanced equation is given in the problem.

Calculate the initial amounts:

\[
\begin{align*}
1.50 \times 10^3 \text{ kg Propene} & \quad \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol}}{42.08 \text{ g}} \right) = 35.6 \times 10^3 \text{ mol} \\
6.80 \times 10^2 \text{ kg NH}_3 & \quad \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol}}{17.04 \text{ g}} \right) = 39.9 \times 10^3 \text{ mol} \\
1.92 \times 10^3 \text{ kg O}_2 & \quad \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol}}{32.00 \text{ g}} \right) = 60.0 \times 10^3 \text{ mol}
\end{align*}
\]

Next, construct an amounts table:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>2 C_3H_6</th>
<th>+</th>
<th>2 NH_3</th>
<th>+</th>
<th>3 O_2</th>
<th>→</th>
<th>2 C_3H_3N</th>
<th>+</th>
<th>6 H_2O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial amount (10³ mol)</td>
<td>35.6</td>
<td>39.9</td>
<td>60.0</td>
<td>0</td>
<td>0</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>kmol/coef</td>
<td>17.8 (LR)</td>
<td>19.95</td>
<td>20.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change (10³ mol)</td>
<td>−35.6</td>
<td>−35.6</td>
<td>3/2</td>
<td>(35.6)</td>
<td>+35.6</td>
<td>+3(35.6)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final amount (10³ mol)</td>
<td>0.0</td>
<td>4.3</td>
<td>6.6</td>
<td>35.6</td>
<td>106.8</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Do a mole–mass conversion to determine the mass of C,H,N that could be produced:

\[
35.6 \times 10^3 \text{ mol} \left( \frac{53.06 \text{ g}}{1 \text{ mol}} \right) \left( \frac{10^{-3} \text{ kg}}{1 \text{ g}} \right) = 1.89 \times 10^3 \text{ kg}
\]

3. (2 points)
Elemental phosphorus, $P_4$, reacts vigorously with oxygen to give $P_4O_{10}$. Use a table of amounts to determine how much $P_4O_{10}$ can be prepared from 3.75 g of $P_4$ and 6.55 g of $O_2$, and how much of the excess reactant will remain at the end of the reaction.

The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the masses of the product formed and the remaining reactant.

Begin by determining the balanced chemical equation:

$$P_4 + O_2 \rightarrow P_4O_{10}$$

$P$ is already balanced. To balance $O$, give $O_2$ a coefficient of 5:

$$P_4 + 5 O_2 \rightarrow P_4O_{10}$$

Next, calculate the initial amounts:

$$3.75 \text{ g } P_4 \times \left( \frac{1 \text{ mol}}{123.9 \text{ g}} \right) = 0.0303 \text{ mol} \quad 6.55 \text{ g } O_2 \times \left( \frac{1 \text{ mol}}{32.00 \text{ g}} \right) = 0.205 \text{ mol}$$

Construct an amounts table:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$P_4$</th>
<th>$5 O_2$</th>
<th>$P_4O_{10}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial amount (mol)</td>
<td>0.0303</td>
<td>0.205</td>
<td>0</td>
</tr>
<tr>
<td>mol/coeff</td>
<td>0.0303 (LR)</td>
<td>0.0410</td>
<td></td>
</tr>
<tr>
<td>Change (mol)</td>
<td>$-0.0303$</td>
<td>$-5(0.0303)$</td>
<td>$+0.0303$</td>
</tr>
<tr>
<td>Final amount (mol)</td>
<td>0</td>
<td>0.0535</td>
<td>0.0303</td>
</tr>
</tbody>
</table>

Now obtain the mass of $P_4O_{10}$ produced, using the information from the amounts table and the molar mass: $M = 4(30.97 \text{ g/mol}) + 10(16.00 \text{ g/mol}) = 283.9 \text{ g/mol}$

The mass that could be produced is $0.0303 \text{ mol} \times \left( \frac{283.9 \text{ g}}{1 \text{ mol}} \right) = 8.60 \text{ g } P_4O_{10}$

Finally, determine the mass of $O_2$ left over:

There would be $0.0535 \text{ mol} \times \left( \frac{32.00 \text{ g}}{1 \text{ mol}} \right) = 1.71 \text{ g } O_2$ left unreacted.

4. (1 points)

Calculate the mass of lithium that has the same number of atoms as 5.75 g of platinum.

“Same number of atoms” also means “same number of moles,” so work with moles:
Concentrated ammonia is 14.8 M. What volume of this solution should be used to prepare 1.25 L of 0.500 M NH₃?

This is a dilution type problem. Rearrange the dilution equation to give an expression for the initial volume:

\[ c_i V_i = c_f V_f \]

\[ V_i = \frac{c_f V_f}{c_i} \]

\[ V_i = \frac{(0.500 \text{ M})(1.25 \text{ L})}{14.8 \text{ M}} = 0.0422 \text{ L or 42.2 mL} \]

Balance the following redox reactions

(e) \( \text{NO}_3^- + \text{H}_2\text{O}_2 \rightarrow \text{NO} + \text{O}_2 \) \hspace{1cm} \text{(basic)}

(f) \( \text{BrO}_3^- + \text{Fe}^{2+} \rightarrow \text{Br}^- + \text{Fe}^{3+} \) \hspace{1cm} \text{(acidic)}

\( \text{NO}_3^- \rightarrow \text{NO} \)

(1) N is balanced; (2) add 2 H₂O on the right; (3) add 4 H₂O on the left and 4 OH⁻ on the right; (4) add 3 e⁻ on the left:
NO₃⁻ + 4 H₂O + 3 e⁻ → NO + 2 H₂O + 4 OH⁻

Cancel duplicated species: NO₃⁻ + 2 H₂O + 3 e⁻ → NO + 4 OH⁻

H₂O₂ → O₂

(1) No elements except H and O; (2) O is balanced; (3) add 2 H₂O on the right and 2 OH⁻ on the left; (4) add 2 e⁻ on the right: H₂O₂ + 2 OH⁻ → O₂ + 2 H₂O + 2 e⁻

Multiply the first reaction by 2 and the second reaction by 3, and then add:

\[
\frac{2[\text{NO}_3^- + 2 \text{H}_2\text{O} + 3 \text{e}^- \rightarrow \text{NO} + 4 \text{OH}^-] + 3[\text{H}_2\text{O}_2 + 2 \text{OH}^- \rightarrow \text{O}_2 + 2 \text{H}_2\text{O} + 2 \text{e}^-]}{2 \text{NO}_3^- + 3 \text{H}_2\text{O}_2 \rightarrow 2 \text{NO} + 3 \text{O}_2 + 2 \text{H}_2\text{O} + 2 \text{OH}^-}
\]

(f) BrO₃⁻ → Br⁻

(1) Br is balanced; (2) add 3 H₂O on the right; (3) add 6 H₃O⁺ on the left and 6 H₂O on the right; (4) add 6 e⁻ on the left: BrO₃⁻ + 6 H₃O⁺ + 6 e⁻ → Br⁻ + 9 H₂O

Fe²⁺ → Fe³⁺

(1–3) All elements are balanced; (4) add 1 e⁻ on the right: Fe²⁺ → Fe³⁺ + e⁻ Multiply the second reaction by 6 and add:

\[
\frac{[\text{BrO}_3^- + 6 \text{H}_3\text{O}^+ + 6 \text{e}^- \rightarrow \text{Br}^- + 9 \text{H}_2\text{O}] + 6[\text{Fe}^2+ \rightarrow \text{Fe}^3+ + \text{e}^-]}{\text{BrO}_3^- + 6 \text{Fe}^{2+} + 6 \text{H}_3\text{O}^+ \rightarrow \text{Br}^- + 6 \text{Fe}^{3+} + 9 \text{H}_2\text{O}}
\]

7. (2 points)
A mouse is placed in a sealed chamber filled with air at 765 Torr and equipped with enough solid KOH to absorb any CO₂ and H₂O produced. The gas volume in the chamber is 2.05 L, and its temperature is held at 298 K. After 2 hours, the pressure inside the chamber has fallen to 725 Torr. What mass of oxygen has the mouse consumed?

As the mouse breathes, it inhales air containing oxygen. Some of this oxygen is used for
metabolism, resulting in CO₂, which the mouse exhales. The amount of oxygen consumed is therefore equal to the amount of CO₂ exhaled. The solid KOH in the chamber absorbs all the CO₂ and any water the mouse exhales. Any reduction in pressure is therefore due to the removal of oxygen from the atmosphere.

The initial amount of gas in the chamber is

\[
\frac{n_{\text{initial}}}{RT} = \frac{p_{\text{initial}}V}{750.06 \text{ Torr}} = \frac{1 \text{ bar}}{(0.08314 \text{ L bar} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298 \text{ K})} = 0.08439 \text{ mol}
\]

And after two hours, the amount of gas remaining is

\[
\frac{n_{\text{final}}}{RT} = \frac{p_{\text{final}}V}{750.06 \text{ Torr}} = \frac{1 \text{ bar}}{(0.08314 \text{ L bar} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298 \text{ K})} = 0.07998 \text{ mol}
\]

The amount of oxygen consumed is therefore \(0.08439 - 0.07998 \text{ mol} = 0.00441 \text{ mol}\).

\((0.00441 \text{ mol})(32.00 \text{ g/mol}) = 0.141 \text{ g O}_2.\)

8. (3 points)
The methanation reaction, \(3 \text{ H}_2 + \text{ CO} \rightarrow \text{ CH}_4 + \text{ H}_2\text{O}\), is used commercially to prepare methane. A gas reactor with a volume of \(1.00 \times 10^2 \text{ L}\) is pressurized at 575 K with 20.0 bar of hydrogen gas and 10.0 bar of CO gas; 145 g of CH₄ is produced. What is the percent yield of the synthesis?

In addition to a yield problem, information is given about both starting materials, so this is a limiting reactant problem that involves gases. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the theoretical yield. Water can be ignored, because CH₄ is the product of interest. All of the reagents of interest are gases, so pressures can be used as the measures of amounts:

<table>
<thead>
<tr>
<th>Reaction:</th>
<th>3 H₂ (g)</th>
<th>+</th>
<th>CO (g)</th>
<th>→</th>
<th>CH₄ (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial pressure (bar)</td>
<td>20.0</td>
<td>10.0</td>
<td></td>
<td></td>
<td>0.0</td>
</tr>
<tr>
<td>Change (bar)</td>
<td>-20.0</td>
<td>-6.67</td>
<td></td>
<td></td>
<td>6.67</td>
</tr>
<tr>
<td>Final pressure (bar)</td>
<td>0.0</td>
<td>3.33</td>
<td></td>
<td></td>
<td>6.67</td>
</tr>
</tbody>
</table>
Use the ideal gas equation and mole–mass conversion to obtain the theoretical yield in grams:

\[
n = \frac{pV}{RT} = \frac{(6.67 \text{ bar})(100 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(575 \text{ K})} = 14.0 \text{ mol}
\]

Theoretical amount = \( nM = (14.0 \text{ mol})(16.04 \text{ g/mol}) = 225 \text{ g} \)

Obtain the percent yield by dividing the actual yield by the theoretical yield:

\[
\% \text{ Yield} = 100\% \left( \frac{145 \text{ g}}{225 \text{ g}} \right) = 64.4\%
\]

9. (4 points)

Two chambers are connected by a valve. One chamber has a volume of 15 L and contains \( N_2 \) gas at a pressure of 2.0 bar. The other has a volume of 1.5 L and contains \( O_2 \) gas at 3.0 bar. The valve is opened, and the two gases are allowed to mix thoroughly. The temperature is constant at 300 K throughout this process. (a) How many moles each of \( N_2 \) and \( O_2 \) are present? (b) What are the final pressures of \( N_2 \) and \( O_2 \), and what is the total pressure?

(a) To determine the number of moles of each gas, we only need the information of the individual containers. Moles can be calculated by direct application of the ideal gas equation:

\[
n_{N_2} = \frac{pV}{RT} = \frac{(2.0 \text{ bar})(15 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 1.2 \text{ mol}
\]

\[
n_{O_2} = \frac{pV}{RT} = \frac{(3.0 \text{ bar})(1.5 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 0.18 \text{ mol}
\]

(b) Remember that the pressure of one gas will not affect the pressure of the other. Therefore, treat the expansion as a change of volume at constant \( n \) and \( T \), so \( pV = \) constant, thus \( pV_f = pV_i \)

\[
p_{f, N_2} = \frac{p_i V_i}{V_f} = \frac{(2.0 \text{ bar})(15 \text{ L})}{(15+1.5 \text{ L})} = 1.8 \text{ bar}
\]

\[
p_{f, O_2} = \frac{p_i V_i}{V_f} = \frac{(3.0 \text{ bar})(1.5 \text{ L})}{(15+1.5 \text{ L})} = 0.27 \text{ bar}
\]

10. (2 points)
A sample of gas is found to exert a pressure of 322 Torr when it is in a 2.00 L flask at 100.00 °C. Compute (a) the new volume if P becomes 525 Torr and T is unchanged; (b) the new pressure if the volume is reduced to 1.50 L and T becomes 50.0 °C; and (c) the new pressure if half the gas is removed but V and T remain the same.

Each part of this problem represents a change of one or more conditions, for which a rearranged version of the ideal gas law can be used.

(a) $T$ and $n$ are fixed, so $pV = nRT = \text{constant}$ and $V_f = \frac{p_i V_i}{p_f}$

$$V_f = \frac{(322 \text{ Torr})(2.00 \text{ L})}{525 \text{ Torr}} = 1.23 \text{ L}$$

(b) Here, pressure, temperature, and volume are changing. So the rearranged ideal gas law is

$$P_f = \frac{p_i V_i T_i}{V_f T_i} = \frac{(322 \text{ Torr})(2.00 \text{ L})(50.0 + 273.15 \text{ K})}{(1.50 \text{ L})(100.00 + 273.15 \text{ K})} = 372 \text{ Torr}$$

(c) Here, only pressure and number of moles are changing. So the ideal gas law can be rearranged to $p_f = \frac{p_i n_f}{n_i}$

$$P_f = \frac{(322 \text{ Torr})(1 \text{ mol})}{2 \text{ mol}} = 161 \text{ Torr}$$

11. (3 points)

A 9.50 g copper block, initially at 200.0 °C, is dropped into a Thermos flask containing 200 mL of water initially at 5.00 °C. What is the final temperature of the Thermos flask contents? $C_{\text{Cu}} = 0.385 \text{ J/g °C}$ (molar 24.435); $C_{\text{H}_2\text{O}} = 4.18 \text{ J/g °C}$ (molar 75.291)

To work a problem involving heat transfers, it is useful to set up a block diagram illustrating the process. In this problem, a copper block transfers energy to water:
Thus, \( q_{\text{water}} = -q_{\text{Cu}} \), \( q_{\text{Cu}} = (nC\Delta T)_{\text{Cu}} \), and \( q_{\text{water}} = (nC\Delta T)_{\text{water}} \)

For Cu, \( n = \frac{m}{M} = \frac{9.50\ g}{63.546\ g/mol} = 0.1495\ \text{mol Cu} \)

\[ C = 24.435\ \text{J mol}^{-1}\ \text{oC}^{-1} \quad \Delta T = x - 200.0\ \text{oC} \]

For water, \( n = \frac{m}{M} = 200\ \text{mL} \left( \frac{1.00\ g}{1\ mL} \right) \left( \frac{1\ mol}{18.02\ g} \right) = 11.10\ \text{mol water} \)

\[ C = 75.291\ \text{J mol}^{-1}\ \text{oC}^{-1} \quad \Delta T = (x - 5.00)\ \text{oC} \]

Substitute and solve for \( x \):

\[(11.10\ \text{mol})(75.291\ \text{J mol}^{-1}\ \text{oC}^{-1})(x - 5.00\ \text{oC}) = -(0.1495\ \text{mol})(24.435\ \text{J mol}^{-1}\ \text{oC}^{-1})(x - 200.0\ \text{oC})\]

\[835.7\ x - 4179 = -(3.653\ x - 730.6)\]

\[835.7\ x - 4179 = -3.653\ x + 730.6\]

\[839.4\ x = 4910\]

\[x = 5.85\ \text{oC} \]

The final temperature of both the copper block and the water is 5.85 °C.

**12. (5 points)**

A sample of carbon monoxide gas with mass of 7.75 g is heated from 25 °C to 175 °C at constant pressure of 5.00 bar. Calculate \( q \), \( w \), and \( E \) for the process. (Be aware of the appropriate units! For \( w \) (L bar) x 101.325 = joules)

\( \Delta E = q + w \quad q = nC\Delta T \quad w_{\text{sys}} = -p\Delta V_{\text{sys}} \quad C_{\text{CO}} = 29.142\ \text{J mol}^{-1}\ \text{oC}^{-1} \)

A diagram helps us visualize what must happen in the process described in this problem:

```
CO gas
m = 7.75 g
T = 25 °C
p = 5.00 bar

heat

q

CO gas
m = 7.75 g
T = 175 °C
p = 5.00 bar
```

The process is heating at constant pressure. The volume will increase during this heating, so work is done as well as heat being added. We have to use Equations 3-3, 3-2, and 3-8 to do the calculation of \( \Delta E \):
\[ \Delta E = q + w \quad q = nC\Delta T \quad w_{sys} = -p\Delta V_{sys} \]

To calculate \( q \), we need moles and heat capacity of CO:

\[
\begin{align*}
n &= \frac{m}{M} = \frac{7.75 \text{ g}}{28.01 \text{ g/mol}} = 0.2767 \text{ mol} \\
&= 175 - 25 = 150. \degree C
\end{align*}
\]

\[
V = \frac{nRT}{p}
\]

Initial: \( V = \frac{(0.2767 \text{ mol})(0.08206 \text{ L bar mol}^{-1} \text{ K}^{-1})(25 + 273 \text{ K})}{5.00 \text{ bar}} = 1.353 \text{ L} \)

Final: \( V = \frac{(0.2767 \text{ mol})(0.08206 \text{ L bar mol}^{-1} \text{ K}^{-1})(175 + 273 \text{ K})}{5.00 \text{ bar}} = 2.034 \text{ L} \)

\[
q = (0.2767 \text{ mol})(29.142 \text{ J mol}^{-1} \degree \text{C}^{-1})(150. \degree C) = 1209 \text{ J}
\]

\[
w = -p\Delta V_{sys} = -(5.00 \text{ bar})(2.034 \text{ L} - 1.353 \text{ L}) = 3.405 \text{ L bar}
\]

\[
\text{Convert } w \text{ to joules: } w = 3.405 \text{ L bar } \left( \frac{101.325 \text{ J}}{1 \text{ L bar}} \right) = -345 \text{ J}
\]

\[
\Delta E = q + w = 1209 \text{ J} - 345 \text{ J} = 864 \text{ J}
\]

**13. (2 points)**

The most common form of elemental sulphur is \( S_8 \), in which eight sulphur atoms link in a ring of single bonds. At high temperature, in the gas phase, \( S_8 \) can break apart to give \( S_2 \), the sulphur analogue of molecular oxygen:

\[ S_8 (g) \rightarrow 4S_2 (g) \quad \Delta H = +219 \text{ kJ/mol at } T = 800 \text{ K} \]

Calculate the \( S=S \) double-bond energy in \( S_2(g) \). \( \text{BE}_{S,S} = 240 \text{kJ/mol} \)

To work this problem, we first draw the structures of the reactant and product, from which we can count the number of bonds of each type.
Reactants: 8 S-S single bonds  Products: 4 S=S bonds

\[ \Delta H_{\text{reaction}} = \Sigma B_H \text{bonds broken} - \Sigma B_H \text{bonds formed} = 8 \text{BE (S-S)} - 4 \text{BE (S=)} \]

\[ 4 \text{BE (S=)} = 8 \text{BE (S-S)} - \Delta H_{\text{reaction}} \]

\[ \text{BE (S=S)} = \frac{(8 \text{ mol})(240 \text{ kJ/mol}) - 219 \text{ kJ}}{4 \text{ mol}} = 425 \text{ kJ/mol} \]

14. (2 points)
Determine the frequencies that hydrogen atoms emit in transitions from the \( n = 6 \) and \( n = 5 \) levels to the \( n = 3 \) level. \( E_n = (-2.18 \times 10^{-18} \text{ J})/n^2 \)

Energies and frequencies for transitions in hydrogen atoms can be calculated from the equation for hydrogen atom energy levels:

\[ E_n = \frac{-2.18 \times 10^{-18} \text{ J}}{n^2} \]

\[ \Delta E_{6-3} = E_6 - E_3 = \left(\frac{-2.18 \times 10^{-18} \text{ J}}{6^2}\right) - \left(\frac{-2.18 \times 10^{-18} \text{ J}}{3^2}\right) = 1.817 \times 10^{-19} \text{ J} \]

\[ v = \frac{E}{h} = \frac{1.817 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 2.74 \times 10^{14} \text{ s}^{-1} \]

\[ \Delta E_{5-3} = E_5 - E_3 = \left(\frac{-2.18 \times 10^{-18} \text{ J}}{5^2}\right) - \left(\frac{-2.18 \times 10^{-18} \text{ J}}{3^2}\right) = 1.550 \times 10^{-19} \text{ J} \]

\[ v = \frac{E}{h} = \frac{1.550 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 2.34 \times 10^{14} \text{ s}^{-1} \]
15. (1 points)
List the values for the quantum numbers for a 3d electron that has spin up and the largest possible value for its magnetic quantum number.

The designation 3d specifies that \( n = 3 \) and \( l = 2 \); the largest possible value for \( m_l \) is \(+l\), in this case \(+2\); and “spin up” means \( m_s = + \frac{1}{2} \). Thus, the values are \( 3, 2, 2, + \frac{1}{2} \).

16. (2 points)
Draw Lewis structures and ball-and-stick structures (or broken lines and solid lines) showing the geometries of molecules of the following substances: (a) methyl isocyanate (CH\(_3\)NCO, a toxic compound responsible for thousands of deaths in Bhopal, India, in 1984); and (b) tetrafluorohydrazine (N\(_2\)F\(_4\)), a colourless liquid used as rocket fuel).

Determine Lewis structures following the standard procedures. Draw ball-and-stick models that show the molecular geometries as determined by the steric numbers of the inner atoms.

(a) CH\(_3\)NCO has \( 4 + 3(1) + 5 + 4 + 6 = 22 \) valence electrons. The bonding framework requires six pairs, and three pairs go on the outer oxygen atom. This leaves four electrons, which are placed on nitrogen, the more electronegative inner atom:

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{N} \quad \text{C} \quad \text{O} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

Complete the octet on the C atom by moving one lone pair each from N and O to form two double bonds:

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{N} \quad \text{C} \quad \text{O} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

The methyl C has \( SN = 4 \) and tetrahedral geometry, N has \( SN = 3 \) with one lone pair for a bent geometry, and the isocyanate C has \( SN = 2 \) and linear geometry:
(b) $\text{N}_2\text{F}_4$ has $2(5) + 4(7) = 38$ valence electrons. Five pairs are used in the bonding framework and 12 pairs are placed around the fluorine atoms, leaving four electrons, two of which are placed on each of the nitrogen atoms. $\text{SN}_\text{N} = 4$, indicating tetrahedral electron pair geometry about each nitrogen atom and trigonal pyramidal shape about each nitrogen atom:

17. (2 points)
Benzyne is an unstable molecule that can be generated as a short-lived species in solution. Suggest a reason why benzyne is very reactive. (Hint: Examine the bond angles.)

The six-membered ring of benzyne requires bond angles of $120^\circ$. The two carbon atoms involved in the triple bond have $\text{SN} = 2$, for which the optimum VSEPR angle is $180^\circ$. Thus, these atoms react readily, either by adding another bonded atom to generate $\text{SN} = 3$ or by
breaking the ring so the bond angle can become 180°.

18. (3 points)
Nitrogen molecules can absorb photons to generate excited-state molecules. Construct an energy-level diagram and place the valence electrons so that it describes the most stable excited state of an N₂ molecule. Is the N-N bond in this excited-state N₂ molecule stronger or weaker than the N-N bond in ground-state nitrogen? Explain your answer.

To generate the configuration of an excited state, move an electron from an occupied orbital to an unoccupied orbital. The most stable excited state results from moving an electron from the least stable occupied orbital to the most stable unoccupied orbital:

![Energy-level diagram for N₂ ground and excited states](image)

The excited state has one more antibonding electron and one less bonding electron than the ground state, so the excited state has a weaker N–N bond than the ground state.

19. (1 points)
Identify the hybrid orbitals used by the inner atoms.

![Lewis structures](image)

The steric number of an inner atom uniquely determines its hybridization. Use the Lewis structure of the molecule to find steric numbers: (a) SN = 3, sp³ hybrids; (b) SN = 6, sp³d hybrids; (c) SN = 5, sp³d² hybrids.
20. (5 points)
A student proposes the following mechanism for the atmospheric decomposition of ozone to molecular oxygen:

\[
\begin{align*}
  O_3 + O_2 & \rightleftharpoons O_5 \quad \text{(fast, reversible)} \\
  O_5 & \rightarrow 2 O_2 + O \quad \text{(slow, rate-determining)}
\end{align*}
\]

(a) Propose a third step that completes this mechanism. (b) Determine the rate law predicted by this mechanism.

(a) The steps of a mechanism must sum to give the observed overall stoichiometry of the reaction. For ozone decomposition, this is \( 2 O_3 \rightarrow 3 O_2 \). The two steps proposed by the student consume 1 \( O_3 \), produce 1 \( O_2 \), and generate an \( O \) atom, which must be consumed. Thus, the third step is \( O_3 + O \rightarrow 2 O_2 \).

(b) The rate law is determined by the rate-determining step: \( \text{Rate} = k_2 [O_5] \). This is not satisfactory, however, because it contains the concentration of an intermediate. Set the rates equal for the forward and reverse first step:

\[
k_1 [O_3][O_2] = k_{-1} [O_5]
\]

Solve this equality for \([O_5]\):

\[
[O_5] = \left( \frac{k_1}{k_{-1}} \right) [O_3][O_2]
\]

Substitute into the rate expression:

\[
\text{Rate} = k_2 \left( \frac{k_1}{k_{-1}} \right) [O_3][O_2]
\]
21. (4 points)
The conversion of C₅H₁₁Br into C₅H₁₀ follows first-order kinetics, with a rate constant of 0.385 h⁻¹. If the initial concentration of C₅H₁₁Br is 0.125 M, find (a) the time at which the concentration will be 1.25 × 10⁻³ M; and (b) the concentration after 3.5 h of reaction.

This is stated to be a first-order reaction, so \( \text{rate} = k[A_{C₅H₁₁Br}] \) and Equation 13-3 applies:

\[
kt = \ln \left( \frac{[A]_n}{[A]} \right)
\]

(a) \( kt = \ln \left( \frac{0.125}{1.25 \times 10^{-3}} \right) = 4.61 \)

\[
(3.5 \text{ h})(0.385 \text{ h}^{-1}) = 12.0 \text{ h}
\]

b) \( \ln[A] = \ln[0.125] - 3.43 \)

\[
[A] = e^{-3.43} = 0.0324 \text{ M or } 3.24 \times 10^{-2} \text{ M}
\]

22. (2 points)
The equilibrium constant for the following reaction is 1.6 × 10⁵ at 1024 K:

\[
\text{H}_2(g) + \text{Br}_2(g) \rightleftharpoons 2\text{HBr}(g)
\]

Find the equilibrium pressures of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K

To calculate concentrations at equilibrium from initial conditions, set up a concentration table. For a gas-phase reaction, concentrations must be expressed in bar. Let \( x \) = change in \( p_{H_2} \):

<table>
<thead>
<tr>
<th>Reaction:</th>
<th>H₂</th>
<th>Br₂</th>
<th>2 HBr</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial pressure (bar)</td>
<td>0</td>
<td>0</td>
<td>10.0</td>
</tr>
<tr>
<td>Change in pressure (bar)</td>
<td>+x</td>
<td>+x</td>
<td>-2x</td>
</tr>
<tr>
<td>Equilibrium pressure (bar)</td>
<td>x</td>
<td>x</td>
<td>10.0 - 2x</td>
</tr>
</tbody>
</table>
Now substitute into the equilibrium constant expression and solve for $x$:

$$K_{eq} = \frac{(p_{HBr})_{eq}^2}{(p_{H_2})_{eq}(p_{Br_2})_{eq}} = \frac{(10.0 - 2x)^2}{(x)(x)} = 1.6 \times 10^5$$

To simplify, assume that $2x \ll 10.0$:

$$1.6 \times 10^5 = \frac{(10.0)^2}{(x)^2} \quad \text{so} \quad x^2 = \frac{100}{1.6 \times 10^5} = 6.25 \times 10^{-4}$$

$$x = 2.5 \times 10^{-2} = (p_{H_2})_{eq} = (p_{Br_2})_{eq}$$

$$(p_{HBr})_{eq} = 10.0 - (2)(2.5 \times 10^{-2}) = 10.0 \text{ bar}$$

$$2(2.5 \times 10^{-2}) = 0.050 \ll 10.0, \text{ so the approximation is valid.}$$

23. (4 points)

$K_{eq}$ for the Haber reaction is $2.81 \times 10^{-5}$ at 472 °C. If a reaction starts with 3.0 bar of $H_2$ and 5.0 bar of $N_2$ at 472 °C, what is the equilibrium pressure of $NH_3$?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

This problem describes an equilibrium reaction. We are asked to determine the equilibrium pressures of all the gases. To calculate pressures at equilibrium from initial conditions, set up a concentration table, write the $K$ expression, and solve for the pressures:

Let $-x = \text{change in } [N_2]$:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$N_2$</th>
<th>$3H_2$</th>
<th>$2NH_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial pressure (bar)</td>
<td>5.0</td>
<td>3.0</td>
<td>0</td>
</tr>
<tr>
<td>Change in pressure (bar)</td>
<td>$-x$</td>
<td>$-3x$</td>
<td>$+2x$</td>
</tr>
<tr>
<td>Equilibrium pressure (bar)</td>
<td>$5.0 - x$</td>
<td>$3.0 - 3x$</td>
<td>$2x$</td>
</tr>
</tbody>
</table>

Now substitute into the equilibrium constant expression and solve for $x$:
\[ K_{\text{eq}} = 2.81 \times 10^{-5} = \frac{(p_{\text{NH}_3})^2}{(p_{\text{N}_2})(p_{\text{H}_2})^2} = \frac{(2x)^2}{(5.0-x)(3.0-3x)^3}; \text{ assume that } 3x << 3.0: \]

\[ 2.81 \times 10^{-5} = \frac{4x^2}{(5.0)(3.0)^3} \quad 4x^2 = (5.0)(3.0)^3(2.81 \times 10^{-5}) = 3.79 \times 10^{-3} \]

\[ x = 3.08 \times 10^{-2} \quad (p_{\text{NH}_3})_{\text{eq}} = 2(3.08 \times 10^{-2}) = 6.2 \times 10^{-2} \text{ bar} \]

\[ 3(3.08 \times 10^{-2}) = 0.09, \text{ about } 3\% \text{ of } 3.0, \text{ so } 3x \ll 3.0 \text{ and the approximation is valid.} \]

24. (4 points)

At 1020 °C, \( K_p = 167.5 \) for the conversion of \( \text{CO}_2(\text{g}) \) to \( \text{CO}(\text{g}) \) by solid graphite, \( \text{C}(\text{s}) \). A 1.00 L, high-pressure chamber containing excess graphite powder is charged with 0.494 mole each of \( \text{CO}_2 \) and \( \text{CO} \) and then is heated to 1020 °C. What is the equilibrium total pressure?

To find an equilibrium total pressure, it is necessary to calculate equilibrium partial pressures of all gaseous participants. First determine the initial pressures of the gases, using the ideal gas equation:

\[ p = \frac{nRT}{V} = \frac{(0.494 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(1020 + 273 \text{ K})}{1.00 \text{ L}} = 53.1 \text{ bar} \]

Set up a concentration table, write the equilibrium expression and solve for the pressure:

<table>
<thead>
<tr>
<th>Reaction:</th>
<th>C (s) +</th>
<th>CO(_2) (g)</th>
<th>2 CO (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial pressure (bar)</td>
<td>( _ _ _ _ )</td>
<td>53.1</td>
<td>53.1</td>
</tr>
<tr>
<td>Change in pressure (bar)</td>
<td>( _ _ _ _ )</td>
<td>( -x )</td>
<td>( +2x )</td>
</tr>
<tr>
<td>Equilibrium pressure (bar)</td>
<td>( _ _ _ _ )</td>
<td>53.1 (-x)</td>
<td>53.1 (+2x)</td>
</tr>
</tbody>
</table>

Substitute into the equilibrium constant expression and solve for \( x \):

\[ 167.5 = \frac{(53.1 + 2x)^2}{(53.1-x)} \quad (167.5)(53.1-x) = (53.1+2x)^2 \]

\[ 8894-167.5x = 2820 + 212.4x + 4x^2 \quad 4x^2 + 379.9x - 6074 = 0 \]

\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-379.9 \pm \sqrt{(379.9)^2 - 4(4)(-6074)}}{2(4)} = \frac{-379.9 \pm 491.4}{8} = 13.9 \]

Rule out the negative value, which would give a negative pressure:
(\(P_{\text{CO}_2}\)\)\text{eq} = 53.1 - 13.9 = 39.2 \text{ bar} \quad (\(P_{\text{CO}}\)\text{eq} = 53.1 + 2(13.9) = 80.9 \text{ bar} \\
p_{\text{total}} = 39.2 + 80.9 = 1.20 \times 10^2 \text{ bar}

25. (6 points)
Calculate the pH of a 2.5 \times 10^{-2} \text{ M solution of the following compounds: (a) NH}_3; (b) HClO; and (c) HCN. \(K_b(\text{NH}_3) = 1.8 \times 10^{-5}\) \(K_a(\text{HClO}) = 4.0 \times 10^{-8}\) \(K_a(\text{HCN}) = 6.2 \times 10^{-10}\)

To calculate the pH of a solution, it is necessary to determine either the hydronium ion concentration or the hydroxide ion concentration.

Determine the nature and initial concentration of each species, construct a concentration table, write the equilibrium expression, and solve for the concentrations:

(a) Weak base, carry out an equilibrium calculation to determine \([\text{OH}^-]\):

<table>
<thead>
<tr>
<th>Reaction: H(_2)O +</th>
<th>NH(_3)</th>
<th>OH(^-) +</th>
<th>NH(_4^+)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial concentration (M)</td>
<td>2.5 \times 10^{-2}</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change in concentration (M)</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Final concentration (M)</td>
<td>2.5 \times 10^{-2} - x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\(K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+]_{\text{eq}}[\text{OH}^-]_{\text{eq}}}{[\text{NH}_3]_{\text{eq}}} = \frac{x^2}{2.5 \times 10^{-2} - x}\); assume that \(x << 2.5 \times 10^{-2}\):

\(1.8 \times 10^{-5} = \frac{x^2}{2.5 \times 10^{-2}}, \text{ so } x = 6.7 \times 10^{-4} \text{ M}; \text{ the assumption is valid.} \]

\([\text{OH}^-] = 6.7 \times 10^{-4} \text{ M} \quad \text{pOH} = -\log(6.7 \times 10^{-4}) = 3.17 \quad \text{pH} = 10.83\)

(b) Weak acid, carry out an equilibrium calculation to determine \([\text{H}_3\text{O}^+]\):

<table>
<thead>
<tr>
<th>Reaction: H(_2)O +</th>
<th>HClO</th>
<th>ClO(^-) +</th>
<th>H(_3)O(^+)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial concentration (M)</td>
<td>2.5 \times 10^{-2}</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change in concentration (M)</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
</tbody>
</table>

\(\text{pH} = 10.83\)
Final concentration (M) | \(2.5 \times 10^{-2} - x\) | \(x\) | \(x\)

\[K_a = 4.0 \times 10^{-8} = \frac{[\text{H}_2\text{O}^+]_{\text{eq}}[\text{HCN}]_{\text{eq}}}{[\text{HCN}]_{\text{eq}}} = \frac{x^2}{2.5 \times 10^{-2} - x}; \text{ assume that } x << 2.5 \times 10^{-2} :\]

\[4.0 \times 10^{-8} = \frac{x^2}{2.5 \times 10^{-2}}\]

\[x^2 = 1.0 \times 10^{-9}, \text{ so } x = 3.2 \times 10^{-5}; \text{ the assumption is valid.}\]

\[[\text{H}_3\text{O}^+] = 3.2 \times 10^{-5} \text{ M} \quad \text{pH} = -\log(3.2 \times 10^{-5}) = 4.49\]

(c) Weak acid, carry out an equilibrium calculation to determine \([\text{H}_3\text{O}^+]:\)

\[
\begin{array}{|c|c|c|c|}
\hline
\text{Reaction: } \text{H}_2\text{O} & \text{HCN} & \text{CN}^- & \text{H}_3\text{O}^+ \\
\hline
\text{Initial concentration (M)} & 2.5 \times 10^{-2} & 0 & 0 \\
\text{Change in concentration (M)} & -x & +x & +x \\
\text{Final concentration (M)} & 2.5 \times 10^{-2} - x & x & x \\
\hline
\end{array}
\]

\[K_a = 6.2 \times 10^{-10} = \frac{[\text{CN}^-]_{\text{eq}}[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HCN}]_{\text{eq}}} = \frac{x^2}{2.5 \times 10^{-2} - x}; \text{ assume that } x << 2.5 \times 10^{-2} :\]

\[6.2 \times 10^{-10} = \frac{x^2}{2.5 \times 10^{-2}}\]

\[x^2 = 1.55 \times 10^{-11}, \text{ so } x = 3.9 \times 10^{-6}; \text{ the assumption is valid.}\]

\[[\text{H}_3\text{O}^+] = 3.9 \times 10^{-6} \text{ M} \quad \text{pH} = -\log(3.9 \times 10^{-6}) = 5.41\]

26. (4 points)
For a solution that is 0.0100 M in NH₄NO₃, do the following:
(a) Identify the major species.
(b) Identify the equilibrium that determines the pH.
(c) Calculate the pH.

\[K_b(\text{NH}_3) = 1.8 \times 10^{-5}\]

Follow standard procedures for dealing with equilibrium problems:

(a) The compound is a salt, so the major species are NH₄⁺, NO₃⁻, and H₂O.
(b) The species with acid–base properties are NH₄⁺ (a weak acid) and H₂O, so the dominant equilibrium is H₂O (l) + NH₄⁺ (aq) → NH₃(aq) + H₃O⁺ (aq).
(c) Carry out an equilibrium calculation to determine [H₃O⁺]:
The equilibrium reaction is a weak acid proton transfer. NH₃ is the species resulting from the loss of a proton from NH₄⁺ so use K_b to determine K_a:

\[ K_b = 1.8 \times 10^{-5} \]

\[ K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} \]

\[ K_x = 5.6 \times 10^{-10} = \frac{[\text{NH}_3]_{\text{eq}}[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{NH}_4^+]_{\text{eq}}} = \frac{x^2}{0.0100 - x}; \text{ assume that } x << 0.0100: \]
\[ 5.6 \times 10^{-10} = \frac{x^2}{0.0100} \]

\[ x^2 = 5.6 \times 10^{-12}, \text{ so } x = 2.37 \times 10^{-6}; \text{ the assumption is valid.} \]

\[ [\text{H}_3\text{O}^+] = 2.37 \times 10^{-6} \text{ M} \]

\[ \text{pH} = -\log(2.37 \times 10^{-6}) = 5.63 \]

---

**27. (5 points)**

According to its label, each tablet of Alka-Seltzer contains 1.916 g (0.0228 mol) of sodium hydrogen carbonate (NaHCO₃). Consider an Alka-Seltzer tablet dissolved in 150 mL of water: (a) What equilibrium determines the pH? (b) What is the pH of the solution?

\( pK_a(\text{HCO}_3^-) = 10.33 \quad pK_b(\text{HCO}_3^-) = 7.65 \)

Follow the standard procedure for dealing with equilibrium calculations:

(a) The major species are Na⁺, HCO₃⁻, and H₂O.

There are two equilibria involving major species:

\[ \text{HCO}_3^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3 (aq) + \text{OH}^- (aq) \]

\( pK_{eq} = pK_{a2} = 10.33 \)

\[ \text{HCO}_3^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3 (aq) + \text{OH}^- (aq) \]

\( pK_{bl} = pK_w - pK_{al} = 14.00 - 6.35 = 7.65 \)

The equilibrium with the larger \( K_{eq} \) (smaller \( pK_{eq} \)) dominates, making this solution basic.

(b) Set up a concentration table, solve for hydroxide ion concentration, and then calculate the pH. Determine the initial concentration using standard stoichiometric procedures:

\[ \left[ \text{HCO}_3^- \right] = \frac{n}{V} = \frac{0.0228 \text{ mol}}{0.150 \text{ L}} = 0.152 \text{ M} \]

<table>
<thead>
<tr>
<th>Reaction: H₂O +</th>
<th>HCO₃⁻</th>
<th>H₂CO₃⁺</th>
<th>OH⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial concentration (M)</td>
<td>0.152</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change in concentration (M)</td>
<td>–x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Final concentration (M)</td>
<td>0.152 – x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

Now substitute into the equilibrium constant expression and solve for \( x \):

\( K_{eq} = 10^{-7.65} = 2.2 \times 10^{-8} \)
\[ 2.2 \times 10^{-8} = \frac{[\text{H}_2\text{CO}_3]_{\text{eq}}[\text{OH}^-]_{\text{eq}}}{[\text{HCO}_3^-]_{\text{eq}}} = \frac{x^2}{0.152 - x} \]: assume that \( x \ll 0.152 \):

\( x^2 = 3.3 \times 10^{-9} \), from which \( x = 5.8 \times 10^{-5} \); the assumption is valid.

\( [\text{OH}^-] = 5.8 \times 10^{-5} \text{ M} \)

\( \text{pOH} = -\log (5.8 \times 10^{-5}) = 4.24 \)

\( \text{pH} = 14.00 - 4.24 = 9.76 \)
Moheur's Periodic Table of the Elements

<table>
<thead>
<tr>
<th>Atomic number</th>
<th>Symbol</th>
<th>Element</th>
<th>Period</th>
<th>Group</th>
<th>Block</th>
<th>Electron configuration</th>
<th>Mass number</th>
<th>Most common isotope</th>
<th>Electron configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>Hydrogen</td>
<td>1</td>
<td>1</td>
<td>s</td>
<td>$1s^1$</td>
<td>1.00784</td>
<td>Most abundant</td>
<td>$1s^1$</td>
</tr>
<tr>
<td>2</td>
<td>He</td>
<td>Helium</td>
<td>1</td>
<td>1</td>
<td>s</td>
<td>$1s^2$</td>
<td>4.002603</td>
<td></td>
<td>$1s^2$</td>
</tr>
<tr>
<td>3</td>
<td>Li</td>
<td>Lithium</td>
<td>2</td>
<td>2</td>
<td>s</td>
<td>$1s^22s^1$</td>
<td>6.941</td>
<td></td>
<td>$1s^22s^1$</td>
</tr>
<tr>
<td>4</td>
<td>Be</td>
<td>Beryllium</td>
<td>2</td>
<td>2</td>
<td>s</td>
<td>$1s^22s^22p^1$</td>
<td>9.01224</td>
<td></td>
<td>$1s^22s^22p^1$</td>
</tr>
<tr>
<td>5</td>
<td>B</td>
<td>Boron</td>
<td>2</td>
<td>18</td>
<td>s</td>
<td>$1s^22s^22p^1$</td>
<td>10.811</td>
<td></td>
<td>$1s^22s^22p^1$</td>
</tr>
<tr>
<td>6</td>
<td>C</td>
<td>Carbon</td>
<td>2</td>
<td>16</td>
<td>s</td>
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</tbody>
</table>

Under normal conditions, bold elements correspond to solid states, bold italic elements correspond to liquid states, and normal elements correspond to gaseous states and normal corresponds to syntactic elements.

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Quadratic equation solution formula:

\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \]

**Data For Water**

- Density = 1.00 g/mL (at 25°C)
- \( s = 2.13 \text{ J g}^{-1} \text{ K}^{-1} \) (solid)
- \( s = 4.184 \text{ J g}^{-1} \text{ K}^{-1} \) (liquid)
- \( s = 2.01 \text{ J g}^{-1} \text{ K}^{-1} \) (gas)
- \( \Delta H^\circ_{\text{fus}} = 6.02 \text{ kJ mol}^{-1} \)
- \( \Delta H^\circ_{\text{vap}} = 40.7 \text{ kJ mol}^{-1} \)

**Constants and Conversion Factors**

<table>
<thead>
<tr>
<th>Conversion Factor</th>
<th>Conversion Factor</th>
<th>Conversion Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mmHg = 1 torr</td>
<td>760 mmHg = 1 atm</td>
<td>1 atm = 101,325 kPa</td>
</tr>
<tr>
<td>1 cm³ = 1 mL</td>
<td>1000 mL = 1 L</td>
<td>1000 L = 1 m³</td>
</tr>
</tbody>
</table>

- Avogadro’s Number \( N \) = 6.022x10²³ mol⁻¹
- Boltzmann’s constant \( k \) = 1.38066x10⁻²³ J K⁻¹
- Faraday’s constant \( F \) = 96,485 C mol⁻¹
- Gas constant \( R \) = 8.31451 J K⁻¹ mol⁻¹
- Planck’s constant \( h \) = 6.62608x10⁻³⁴ J s
- Speed of Light \( c \) = 2.99792458x10⁸ m s⁻¹
**Gas Laws**

\[ PV = nRT \]

\[ \frac{P_{V_1}}{T_1} = \frac{P_{V_2}}{T_2} \]

\[ P_T = P_1 + P_2 + P_3 + \ldots \]

\[ d = \frac{m}{V} = \frac{P \cdot MM}{RT} \]

\[ E_k = \frac{1}{2} mv^2 \]

\[ n_{ma} = \sqrt{\frac{3RT}{MM}} \]

\[ \frac{Rate A}{Rate B} = \left( \frac{MM_B}{MM_A} \right)^{1/2} \]

\[ \left( P + \frac{n^2a}{V^2} \right) (V - nb) = nRT \]

**Equilibrium**

\[ K_p = K_c(RT)^{\Delta n} \]

**Thermochemistry**

\[ \Delta U = q + W \]

\[ W_{\text{system}} = -P\Delta V = -\Delta nRT \]

\[ \Delta H = \Delta U + P\Delta V \]

\[ q_p = \Delta U + P\Delta V \]

\[ q = ms\Delta T \]

\[ \Delta H^{\circ}_{\text{rxn}} = \sum n\Delta H^\circ_i(\text{products}) - \sum n\Delta H^\circ_i(\text{reactants}) \]

**The atom**

\[ E = hv \]

\[ c = \nu\lambda \]

\[ E = -B/n^2 \]

**Kinetics**

\[ [A]_t = [A]_o - kt \]

\[ \ln[A]_t = \ln[A]_o - kt \]

\[ \frac{1}{[A]_o} = \frac{1}{[A]_o} + kt \]

\[ k = Ae^{(-E_a/RT)} \]

\[ \ln(k_2/k_1) = \left( -\frac{E_a}{R} \right) \left( \frac{1}{T_2} - \frac{1}{T_1} \right) \]

**Acid/Base**

\[ pOH = -\log[OH^-] \]

\[ pH = -\log[H^+] \]

\[ pH + pOH = 14 \]

\[ K_a \times K_b = K_w \]

\[ pH = pK_a + \log[A^-]/[HA] \]

\[ pH = \frac{pK_{a1} + pK_{a2}}{2} - \ldots \]