### Periodic Table of the Elements

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|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|
| K     | Ca    | Sc    | Ti    | V     | Cr    | Mn    | Fe    | Co    | Ni    | Cu    | Zn    | Ga    | Ge    | As    | Se    | Br    | Kr    |
| 39.10 | 40.08 | 44.96 | 47.90 | 50.94 | 52.00 | 54.94 | 55.85 | 58.93 | 58.69 | 63.55 | 65.39 | 69.72 | 72.59 | 74.92 | 78.96 | 79.90 | 83.80 |

| 37    | 38    | 39    | 40    | 41    | 42    | 43    | 44    | 45    | 46    | 47    | 48    | 49    | 50    | 51    | 52    | 53    | 54    |
|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|
| Rb    | Sr    | Y     | Zr    | Nb    | Mo    | Tc    | Ru    | Rh    | Pd    | Ag    | Cd    | In    | Sn    | Sb    | Te    | I     | Xe    |
| 85.47 | 87.62 | 88.91 | 91.22 | 92.91 | 95.94 | (98)  | 101.1 | 102.91| 106.42| 107.87| 112.41| 114.82| 118.71| 121.75| 127.60| 126.91| 131.29|

| 55    | 56    | 57    | 72    | 73    | 74    | 75    | 76    | 77    | 78    | 79    | 80    | 81    | 82    | 83    | 84    | 85    | 86    |
|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|-------|
| Cs    | Ba    | La    | Hf    | Ta    | W     | Re    | Os    | Ir    | Pt    | Au    | Cd    | In    | Sn    | Sb    | Te    | I     | Xe    |
| 132.91| 137.33| 138.91| 178.49| 180.95| 183.85| 186.21| 190.2 | 192.2 | 195.08| 196.97| 200.59| 204.38| 207.2 | 208.98| (209) | (210) | (222) |

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*Lanthanide Series

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| Ce    | Pr    | Nd    | Pm    | Sm    | Eu    | Gd    | Tb    | Dy    | Ho    | Er    | Tm    | Yb    | Lu    |
| 140.12| 140.91| 144.24| (145) | 150.4 | 151.97| 157.25| 158.93| 162.50| 164.93| 167.26| 168.93| 173.04| 174.97|

†Actinide Series

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</tr>
</tbody>
</table>
ATOMIC STRUCTURE

\[ E = h \nu \quad c = \lambda v \]
\[ \lambda = \frac{h}{mv} \quad p = mv \]
\[ E_n = \frac{-2.178 \times 10^{-18}}{n^2} \text{ joule} \]

EQUILIBRIUM

\[ K_a = \frac{[H^+][A^-]}{[HA]} \]
\[ K_b = \frac{[OH^-][HB^+]}{[B]} \]
\[ K_w = [OH^-][H^+] = 1.0 \times 10^{-14} \text{ @ } 25^\circ C \]
\[ = K_a \times K_b \]
\[ \text{pH} = -\log[H^+], \quad \text{pOH} = -\log[OH^-] \]
\[ 14 = \text{pH} + \text{pOH} \]
\[ \text{pH} = pK_a + \log\frac{[A^-]}{[HA]} \]
\[ \text{pOH} = pK_b + \log\frac{[HB^+]}{[B]} \]
\[ pK_a = -\log K_a, \quad pK_b = -\log K_b \]
\[ K_p = K_c(RT)^\Delta n, \]
where \( \Delta n = \) moles product gas – moles reactant gas

THERMOCHEMISTRY/KINETICS

\[ \Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants} \]
\[ \Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants} \]
\[ \Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants} \]
\[ \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \]
\[ = -RT\ln K = -2.303RT\log K \]
\[ = -n\overline{E}E^\circ \]
\[ \Delta G = \Delta G^\circ + RT\ln Q = \Delta G^\circ + 2.303RT\log Q \]
\[ q = mc\Delta T \]
\[ C_p = \frac{\Delta H}{\Delta T} \]
\[ \ln[A]_f - \ln[A]_0 = -kt \]
\[ \frac{1}{[A]_f} - \frac{1}{[A]_0} = kt \]
\[ \ln k = \frac{-E_a}{RT} + \ln A \]

\[ \begin{align*}
E &= \text{energy} \\
\nu &= \text{velocity} \\
\nu &= \text{frequency} \\
\lambda &= \text{principal quantum number} \\
m &= \text{mass} \\
p &= \text{momentum} \\
\end{align*} \]

Speed of light, \( c = 3.0 \times 10^8 \text{ m s}^{-1} \)

Planck’s constant, \( h = 6.63 \times 10^{-34} \text{ J s} \)

Boltzmann’s constant, \( k = 1.38 \times 10^{-23} \text{ J K}^{-1} \)

Avogadro’s number = \( 6.022 \times 10^{23} \text{ mol}^{-1} \)

Electron charge, \( e = -1.602 \times 10^{-19} \text{ coulomb} \)

1 electron volt per atom = 96.5 kJ mol\(^{-1}\)

---

Equilibrium Constants

\[ \begin{align*}
K_a &= \text{weak acid} \\
K_b &= \text{weak base} \\
K_w &= \text{water} \\
K_p &= \text{gas pressure} \\
K_c &= \text{molar concentrations} \\
\end{align*} \]

\[ \begin{align*}
S^\circ &= \text{standard entropy} \\
H^\circ &= \text{standard enthalpy} \\
G^\circ &= \text{standard free energy} \\
E^\circ &= \text{standard reduction potential} \\
T &= \text{temperature} \\
n &= \text{moles} \\
m &= \text{mass} \\
q &= \text{heat} \\
c &= \text{specific heat capacity} \\
C_p &= \text{molar heat capacity at constant pressure} \\
E_a &= \text{activation energy} \\
k &= \text{rate constant} \\
A &= \text{frequency factor} \\
\end{align*} \]

Faraday’s constant, \( \overline{E} = 96,500 \text{ coulombs per mole of electrons} \)

Gas constant, \( R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1} \)
\[ = 0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} \]
\[ = 8.31 \text{ volt coulomb mol}^{-1} \text{ K}^{-1} \]
GASES, LIQUIDS, AND SOLUTIONS

\[ PV = nRT \]
\[ \left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT \]
\[ P_A = P_{\text{total}} \times X_A, \text{where} \ X_A = \frac{\text{moles} \ A}{\text{total moles}} \]
\[ P_{\text{total}} = P_A + P_B + P_C + \ldots \]
\[ n = \frac{m}{M} \]
\[ K = \frac{C}{273} \]
\[ \frac{PV_1}{T_1} = \frac{PV_2}{T_2} \]
\[ D = \frac{m}{V} \]
\[ u_{\text{rms}} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{M}} \]

KE per molecule = \frac{1}{2} m v^2
KE per mole = \frac{3}{2} RT

\[ \frac{r_1}{r_2} = \frac{[M_2]}{[M_1]} \]
molarity, \( M \) = moles solute per liter solution
molality = moles solute per kilogram solvent
\[ \Delta T_f = iK_f \times \text{molality} \]
\[ \Delta T_b = iK_b \times \text{molality} \]
\[ \pi = iMRT \]
\[ A = abc \]

OXIDATION-REDUCTION; ELECTROCHEMISTRY

\[ Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{where} \ a + b \rightarrow c + d \]
\[ I = \frac{q}{t} \]
\[ E_{\text{cell}} = \frac{RT}{nF} \ln Q = E_{\text{cell}}^o - \frac{0.0592}{n} \log Q \ @ \ 25^\circ C \]
\[ \log K = \frac{nE^o}{0.0592} \]

\( P \) = pressure
\( V \) = volume
\( T \) = temperature
\( n \) = number of moles
\( D \) = density
\( m \) = mass
\( v \) = velocity

\( u_{\text{rms}} \) = root-mean-square speed
\( KE \) = kinetic energy
\( r \) = rate of effusion
\( M \) = molar mass
\( \pi \) = osmotic pressure
\( i \) = van't Hoff factor
\( K_f \) = molal freezing-point depression constant
\( K_b \) = molal boiling-point elevation constant
\( A \) = absorbance
\( a \) = molar absorptivity
\( b \) = path length
\( c \) = concentration
\( Q \) = reaction quotient
\( I \) = current (amperes)
\( q \) = charge (coulombs)
\( t \) = time (seconds)
\( E^o \) = standard reduction potential
\( K \) = equilibrium constant

Gas constant, \( R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1} \)
\[ = 0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} \]
\[ = 8.31 \text{ volt coulomb mol}^{-1} \text{ K}^{-1} \]

Boltzmann's constant, \( k = 1.38 \times 10^{-23} \text{ J K}^{-1} \)
\( K_f \) for \( \text{H}_2\text{O} = 1.86 \text{ K kg mol}^{-1} \)
\( K_b \) for \( \text{H}_2\text{O} = 0.512 \text{ K kg mol}^{-1} \)
1 atm = 760 mm Hg = 760 torr
STP = 0.00°C and 1.0 atm
Faraday's constant, \( F = 96,500 \text{ coulombs per mole of electrons} \)
CHEMISTRY
Section II
(Total time—95 minutes)

Part A
Time—55 minutes
YOU MAY USE YOUR CALCULATOR FOR PART A.

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

Be sure to write all your answers to the questions on the lined pages following each question in the goldenrod booklet. Do NOT write your answers on the lavender insert.

Answer Questions 1, 2, and 3. The Section II score weighting for each question is 20 percent.

1. A sample of solid \( \text{U}_3\text{O}_8 \) is placed in a rigid 1.500 L flask. Chlorine gas, \( \text{Cl}_2(g) \), is added, and the flask is heated to 862\(^\circ\)C. The equation for the reaction that takes place and the equilibrium-constant expression for the reaction are given below.

\[
\text{U}_3\text{O}_8(s) + 3 \text{Cl}_2(g) \rightleftharpoons 3 \text{UO}_2\text{Cl}_2(g) + \text{O}_2(g) \quad \quad K_p = \frac{(p_{\text{UO}_2\text{Cl}_2})^3(p_{\text{O}_2})}{(p_{\text{Cl}_2})^3}
\]

When the system is at equilibrium, the partial pressure of \( \text{Cl}_2(g) \) is 1.007 atm and the partial pressure of \( \text{UO}_2\text{Cl}_2(g) \) is \( 9.734 \times 10^{-4} \) atm.

(a) Calculate the partial pressure of \( \text{O}_2(g) \) at equilibrium at 862\(^\circ\)C.

(b) Calculate the value of the equilibrium constant, \( K_p \), for the system at 862\(^\circ\)C.

(c) Calculate the Gibbs free-energy change, \( \Delta G^\circ \), for the reaction at 862\(^\circ\)C.

(d) State whether the entropy change, \( \Delta S^\circ \), for the reaction at 862\(^\circ\)C is positive, negative, or zero. Justify your answer.

(e) State whether the enthalpy change, \( \Delta H^\circ \), for the reaction at 862\(^\circ\)C is positive, negative, or zero. Justify your answer.

(f) After a certain period of time, 1.000 mol of \( \text{O}_2(g) \) is added to the mixture in the flask. Does the mass of \( \text{U}_3\text{O}_8(s) \) in the flask increase, decrease, or remain the same? Justify your answer.
2. Answer the following problems about gases.

(a) The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ne-20</td>
<td>19.99</td>
</tr>
<tr>
<td>Ne-22</td>
<td>21.99</td>
</tr>
</tbody>
</table>

(i) Using the information above, calculate the percent abundance of each isotope.

(ii) Calculate the number of Ne-22 atoms in a 12.55 g sample of naturally occurring neon.

(b) A major line in the emission spectrum of neon corresponds to a frequency of $4.34 \times 10^{14} \text{ s}^{-1}$. Calculate the wavelength, in nanometers, of light that corresponds to this line.

(c) In the upper atmosphere, ozone molecules decompose as they absorb ultraviolet (UV) radiation, as shown by the equation below. Ozone serves to block harmful ultraviolet radiation that comes from the Sun.

$$\text{O}_3(g) \xrightarrow{\text{UV}} \text{O}_2(g) + \text{O}(g)$$

A molecule of $\text{O}_3(g)$ absorbs a photon with a frequency of $1.00 \times 10^{15} \text{ s}^{-1}$.

(i) How much energy, in joules, does the $\text{O}_3(g)$ molecule absorb per photon?

(ii) The minimum energy needed to break an oxygen-oxygen bond in ozone is 387 kJ mol$^{-1}$. Does a photon with a frequency of $1.00 \times 10^{15} \text{ s}^{-1}$ have enough energy to break this bond? Support your answer with a calculation.
3. In a hydrogen-oxygen fuel cell, energy is produced by the overall reaction represented above.

\[ 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(l) \]

(a) When the fuel cell operates at 25°C and 1.00 atm for 78.0 minutes, 0.0746 mol of \( \text{O}_2(g) \) is consumed. Calculate the volume of \( \text{H}_2(g) \) consumed during the same time period. Express your answer in liters measured at 25°C and 1.00 atm.

(b) Given that the fuel cell reaction takes place in an acidic medium,
   (i) write the two half reactions that occur as the cell operates,
   (ii) identify the half reaction that takes place at the cathode, and
   (iii) determine the value of the standard potential, \( E^\circ \), of the cell.

(c) Calculate the charge, in coulombs, that passes through the cell during the 78.0 minutes of operation as described in part (a).

STOP

If you finish before time is called, you may check your work on this part only. Do not turn to the other part of the test until you are told to do so.
4. For each of the following three reactions, in part (i) write a balanced equation for the reaction and in part (ii) answer the question about the reaction. In part (i), coefficients should be in terms of lowest whole numbers. Assume that solutions are aqueous unless otherwise indicated. Represent substances in solutions as ions if the substances are extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction. You may use the empty space at the bottom of the next page for scratch work, but only equations that are written in the answer boxes provided will be graded.

**EXAMPLE:**
A strip of magnesium metal is added to a solution of silver(I) nitrate.

(i) Balanced equation:
\[ \text{Mg} + 2\text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag} \]

(ii) Which substance is oxidized in the reaction?

**Mg** is oxidized.

(a) Solid ammonium carbonate decomposes as it is heated.

(i) Balanced equation:

(ii) Predict the algebraic sign of \( \Delta S^\circ \) for the reaction. Explain your reasoning.

______________________________________________________________________________________
(b) Chlorine gas, an oxidizing agent, is bubbled into a solution of potassium bromide.

(i) Balanced equation:

(ii) What is the oxidation number of chlorine before the reaction occurs? What is the oxidation number of chlorine after the reaction occurs?

______________________________________________________________________________________
______________________________________________________________________________________

(c) A small piece of sodium is placed in a beaker of distilled water.

(i) Balanced equation:

(ii) The reaction is exothermic, and sometimes small flames are observed as the sodium reacts with the water. Identify the product of the reaction that burns to produce the flames.

______________________________________________________________________________________
______________________________________________________________________________________
Answer Question 5 and Question 6. The Section II score weighting for these questions is 15 percent each.

Your responses to these questions will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

5. Answer the following questions about laboratory situations involving acids, bases, and buffer solutions.

(a) Lactic acid, \( \text{HC}_3\text{H}_5\text{O}_3 \), reacts with water to produce an acidic solution. Shown below are the complete Lewis structures of the reactants.

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

In the space provided above, complete the equation by drawing the complete Lewis structures of the reaction products.

(b) Choosing from the chemicals and equipment listed below, describe how to prepare 100.00 mL of a 1.00 \( M \) aqueous solution of \( \text{NH}_4\text{Cl} \) (molar mass 53.5 g mol\(^{-1}\)). Include specific amounts and equipment where appropriate.

\[
\begin{align*}
\text{NH}_4\text{Cl(s)} &\quad 50 \text{ mL buret} &\quad 100 \text{ mL graduated cylinder} &\quad 100 \text{ mL pipet} \\
\text{Distilled water} &\quad 100 \text{ mL beaker} &\quad 100 \text{ mL volumetric flask} &\quad \text{Balance}
\end{align*}
\]

(c) Two buffer solutions, each containing acetic acid and sodium acetate, are prepared. A student adds 0.10 mol of \( \text{HCl} \) to 1.0 L of each of these buffer solutions and to 1.0 L of distilled water. The table below shows the pH measurements made before and after the 0.10 mol of \( \text{HCl} \) is added.

<table>
<thead>
<tr>
<th></th>
<th>pH Before HCl Added</th>
<th>pH After HCl Added</th>
</tr>
</thead>
<tbody>
<tr>
<td>Distilled water</td>
<td>7.0</td>
<td>1.0</td>
</tr>
<tr>
<td>Buffer 1</td>
<td>4.7</td>
<td>2.7</td>
</tr>
<tr>
<td>Buffer 2</td>
<td>4.7</td>
<td>4.3</td>
</tr>
</tbody>
</table>

(i) Write the balanced net-ionic equation for the reaction that takes place when the \( \text{HCl} \) is added to buffer 1 or buffer 2.
(ii) Explain why the pH of buffer 1 is different from the pH of buffer 2 after 0.10 mol of \( \text{HCl} \) is added.
(iii) Explain why the pH of buffer 1 is the same as the pH of buffer 2 before 0.10 mol of \( \text{HCl} \) is added.

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GO ON TO THE NEXT PAGE.
6. The table above shows the first three ionization energies for atoms of four elements from the third period of the periodic table. The elements are numbered randomly. Use the information in the table to answer the following questions.

(a) Which element is most metallic in character? Explain your reasoning.

(b) Identify element 3. Explain your reasoning.

(c) Write the complete electron configuration for an atom of element 3.

(d) What is the expected oxidation state for the most common ion of element 2?

(e) What is the chemical symbol for element 2?

(f) A neutral atom of which of the four elements has the smallest radius?