Chemistry 231
Exams I
Fall 2014
Oregon State University
Drs. Nafshun, Watson, Nyman, Barth, Burand

Instructions: You should have with you several number two pencils, an eraser, your 3" x 5" note card, a calculator, and your University ID Card. If you have notes or a phone with you, place them in a sealed backpack and place the backpack OUT OF SIGHT or place the notes directly on the table at the front of the room.

Fill in the front page of the Scantron sheet with your last name, first name, middle initial, student identification number, and section number (below). Leave the test form number blank.

Section 001 (MWF 8am with Dr. Nafshun)
Section 002 (MWF 9am with Dr. Nafshun)
Section 003 (MWF 10am with Dr. Nafshun)
Section 004 (MWF 11am with Dr. Watson)
Section 005 (MWF 1pm with Dr. Nyman)
Section 006 (MWF 2pm with Dr. Barth)
Section 007 (MWF 3pm with Dr. Burand)

This exam consists of 32 multiple-choice questions; each has 5 points attached. When you finish this exam, proceed to the proctor. Flash your OSU ID Card and submit your completed Scantron form. You may take your notecard and this exam packet with you.
### TABLE 1.2 SI Prefix Multipliers

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>exa</td>
<td>E</td>
<td>$10^{18}$</td>
</tr>
<tr>
<td>peta</td>
<td>P</td>
<td>$10^{15}$</td>
</tr>
<tr>
<td>tera</td>
<td>T</td>
<td>$10^{12}$</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>$10^9$</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>$10^6$</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>$10^3$</td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>0.1</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>0.01</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>0.001</td>
</tr>
<tr>
<td>micro</td>
<td>μ</td>
<td>$10^{-6}$</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>$10^{-12}$</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>$10^{-15}$</td>
</tr>
<tr>
<td>atto</td>
<td>a</td>
<td>$10^{-18}$</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Conversion</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 inch = 2.54 cm</td>
</tr>
<tr>
<td>1 mile = 1.60934 km</td>
</tr>
<tr>
<td>1000 mL = 1 L</td>
</tr>
<tr>
<td>1 mL = 1 cm³</td>
</tr>
<tr>
<td>1 mole (Nₐ) = 6.022 x 10²³</td>
</tr>
<tr>
<td>c = 3.00 x 10⁸ m/s</td>
</tr>
<tr>
<td>h = 6.626 x 10⁻³⁴ J·s</td>
</tr>
<tr>
<td>1 m = 1 x 10⁹ nm</td>
</tr>
</tbody>
</table>

### 9 Polyatomic Ions

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydroxide</td>
<td>OH⁻</td>
</tr>
<tr>
<td>Cyanide</td>
<td>CN⁻</td>
</tr>
<tr>
<td>Nitrate</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>Acetate</td>
<td>CH₃COO⁻</td>
</tr>
<tr>
<td>Carbonate</td>
<td>CO₃²⁻</td>
</tr>
<tr>
<td>Phosphate</td>
<td>PO₄³⁻</td>
</tr>
<tr>
<td>Hydronium</td>
<td>H₂O⁺</td>
</tr>
<tr>
<td>Ammonium</td>
<td>NH₄⁺</td>
</tr>
<tr>
<td>Sulfate</td>
<td>SO₄²⁻</td>
</tr>
</tbody>
</table>

### Specific Heat (J/g°C)

<table>
<thead>
<tr>
<th>Material</th>
<th>Specific Heat</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O (l)</td>
<td>4.184</td>
</tr>
<tr>
<td>CH₃CH₂OH (l)</td>
<td>2.42</td>
</tr>
<tr>
<td>Pb (s)</td>
<td>0.128</td>
</tr>
<tr>
<td>Au (s)</td>
<td>0.128</td>
</tr>
<tr>
<td>Ag (s)</td>
<td>0.235</td>
</tr>
<tr>
<td>Cu (s)</td>
<td>0.385</td>
</tr>
<tr>
<td>Fe (s)</td>
<td>0.449</td>
</tr>
<tr>
<td>Al (s)</td>
<td>0.903</td>
</tr>
<tr>
<td>Sand</td>
<td>0.84</td>
</tr>
<tr>
<td>Granite</td>
<td>0.790</td>
</tr>
<tr>
<td>Equation</td>
<td>Description</td>
</tr>
<tr>
<td>----------</td>
<td>-------------</td>
</tr>
<tr>
<td>$\lambda = \frac{h}{m\nu}$</td>
<td></td>
</tr>
<tr>
<td>$\Delta E = q + w$</td>
<td>$q = mc\Delta T$</td>
</tr>
<tr>
<td>$h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}$</td>
<td>$\nu = \frac{c}{\lambda}$</td>
</tr>
<tr>
<td>$\frac{1}{\lambda} = R_{H} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$</td>
<td>$R_{H} = 2.180 \times 10^{-18} \text{ J/photons}$</td>
</tr>
<tr>
<td>$1 \text{ J} = 1 \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}$</td>
<td>electron mass = $9.10938 \times 10^{-31} \text{ kg}$</td>
</tr>
<tr>
<td>$\Psi = H\Psi$</td>
<td>$\Delta E = q + w$</td>
</tr>
<tr>
<td>$\Delta H = \Delta E - P\Delta V$</td>
<td>$P(r) = 4 \pi r^2 \psi^2$</td>
</tr>
<tr>
<td>$E_n = -\frac{hcR_{\infty}}{n^2}$</td>
<td>$E_n = -\frac{Rz}{n^2}$</td>
</tr>
</tbody>
</table>
1. A student combusts a compound in a bomb calorimeter, and the temperature of the calorimeter increases from 25.62 °C to 27.42 °C. The calorimeter has a constant of 91.82 J/°C. How many significant figures can the student report for the ΔE in this reaction?

a) 1
b) 2
\[ q = \left( 91.82 \ \text{J/°C} \right) \left( 27.42 \ °C - 25.62 \ °C \right) \]
\[ = \left( 91.82 \ \text{J/°C} \right) \left( 1.80 \ °C \right) \]
\[ = \frac{165.5}{3} \]
d) 3

e) 5

2. A solid cube is measured to be 0.451 inches on each side. The mass of the cube is 12.9 g. Which of the following metals is this cube most likely to be?

a) Yttrium (4.45 g/cm³)  \[ s = 0.451 \ \text{in} \]
\[ = \frac{1.15 \ \text{cm}}{1 \ \text{in}} \]
b) Lead (11.3 g/cm³)  \[ s = 0.451 \ \text{in} \]
\[ = \frac{2.54 \ \text{cm}}{1 \ \text{in}} \]
c) Iridium (22.4 g/cm³)  \[ V = s^3 = (1.15 \ \text{cm})^3 \]
\[ = 1.50 \ \text{cm}^3 \]
d) Tellurium (4.95 g/cm³)  \[ V = \frac{12.9}{8.58} \]
\[ = 1.50 \ \text{cm}^3 \]
e) Niobium (8.58 g/cm³)

3. Which of the following statements about subatomic particles is TRUE?

a) A neutral atom contains the same number of protons and electrons.

b) Protons have about the same mass as electrons.

c) Electrons make up most of the mass of an atom.

d) Protons and neutrons have opposite, but equal in magnitude, charges.

e) Neutrons and electrons are found in the nucleus of an atom.
4. How many moles of Kr are contained in 398 mg of Kr?

a) \(4.75 \times 10^{-3}\) moles Kr

b) 33.4 moles Kr

c) \(2.11 \times 10^{-4}\) moles Kr

d) \(2.99 \times 10^{-3}\) moles Kr

e) \(1.19 \times 10^{-4}\) moles Kr

\[
398 \times 10^{-3} \text{ g Kr} \left( \frac{1 \text{ mol}}{83.8 \text{ g}} \right) = 4.75 \times 10^{-3} \text{ mol Kr}
\]

5. The fictional element Beaverium (Bv) has two stable isotopes: \(^{145}\text{Bv}\) (mass = 144.9362 amu) and \(^{148}\text{Bv}\) (mass = 147.9177 amu). Which of the following could not be the average atomic mass of Beaverium?

a) 145.7731 amu

b) 144.9968 amu

c) 146.3214 amu

d) 147.9882 amu

e) any of these could be the average atomic mass based on the information given

\[
\text{Average atomic mass} = \frac{144.9362 + 147.9177}{2} = 146.42695 \text{ amu}
\]

6. How many total electrons are present in 0.1415 g of oxide ions? (O\(^2-\) each)

a) \(2.23 \times 10^{23}\) electrons

b) \(3.20 \times 10^{22}\) electrons

c) \(4.26 \times 10^{22}\) electrons

d) \(8.84 \times 10^{22}\) electrons

e) \(5.33 \times 10^{22}\) electrons

\[
0.1415 \text{ g } \text{O}^{2-} \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right)
\]

\[
= \frac{0.00884 \text{ mol } \text{O}^{2-}}{6.022 \times 10^{23} \text{ mol}^{-1}} = 5.33 \times 10^{21} \text{ O}^{2-} \text{ ions}
\]

\[
5.33 \times 10^{21} \text{ O}^{2-} \text{ ions} \left( \frac{10 \text{ e}^-}{1 \text{ O}^{2-}} \right) = 5.33 \times 10^{22} \text{ e}^-
\]
7. Which of the following pairs of formulas and names are incorrect?
   a) Fe$_3$(PO$_4$)$_2$; iron (III) phosphate
   b) NH$_4$OH; ammonium hydroxide
   c) CaCl$_2$; calcium chloride
   d) Al(NO$_3$)$_3$·6H$_2$O; aluminum nitrate hexahydrate
   e) TiS$_2$; titanium (IV) sulfide

8. How many grams of water are in 16.00 g of FeSO$_4$ · 4H$_2$O? Molar Mass of FeSO$_4$ · 4H$_2$O = 224.0 g/mol.
   a) 5.149
   b) 4.571
   c) 6.429
   d) 3.502
   e) 4.857

9. Aqueous iron(II) chloride reacts with aqueous sodium phosphate to produce solid iron(II) phosphate and aqueous sodium chloride. Which represents a balanced equation of this?
   a) 3 FeCl$_2$ (aq) + Na$_3$PO$_4$ (aq) → Fe$_3$(PO$_4$)$_2$ (s) + 3 NaCl (aq)
   b) FeCl$_2$ (aq) + Na$_3$PO$_4$ (aq) → Fe$_3$(PO$_4$)$_2$ (s) + NaCl (aq)
   c) FeCl$_3$ (aq) + Na$_3$PO$_4$ (aq) → Fe$_2$PO$_4$ (s) + 3 NaCl (aq)
   d) 3 FeCl$_2$ (aq) + 2 Na$_3$PO$_4$ (aq) → Fe$_3$(PO$_4$)$_2$ (s) + 6 NaCl (aq)
   e) 3 FeCl$_2$ (aq) + Na$_3$PO$_4$ (aq) → 2 Fe$_3$(PO$_4$)$_2$ (s) + 3 NaCl (aq)

\[ 3 \text{FeCl}_2(\text{aq}) + 2 \text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Fe}_3(\text{PO}_4)_2(\text{s}) + 6 \text{NaCl(\text{aq})} \]
10. 8.84 g of a compound is analyzed and found to contain 3.53 g of C, 0.295 g of H, 4.08 g of N, and 0.935 g of O. Which of the following could be the molecular formula for this compound?

a) \( \text{C}_6\text{H}_{10}\text{N}_2\text{O}_3 \)

\[ 3.53 \text{ g C} \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right) = 0.294 \text{ mol C} / 0.0584 \]

b) \( \text{C}_9\text{H}_{18}\text{O}_3 \)

c) \( \text{C}_6\text{H}_{12}\text{N}_6 \)

\[ 0.295 \text{ g H} \left( \frac{1 \text{ mol}}{1.008 \text{ g}} \right) = 0.293 \text{ mol H} / 0.0584 \]

d) \( \text{C}_3\text{H}_{10}\text{N}_3\text{O}_3 \)

\[ 4.08 \text{ g N} \left( \frac{1 \text{ mol}}{14.01 \text{ g}} \right) = 0.291 \text{ mol N} / 0.0584 \]

\[ 0.935 \text{ g O} \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 0.0584 \text{ mol O} / 0.0584 \]

11. Consider the following reaction:

\[ \text{N}_2 \text{ (g) + 3 H}_2 \text{ (g) } \rightarrow 2 \text{ NH}_3 \text{ (g)} \]

What mass of hydrogen gas must be reacted in excess nitrogen gas to give a theoretical yield of 75.0 g of \( \text{NH}_3 \) gas?

a) 6.60 g \( \text{H}_2 \)

b) 13.3 g \( \text{H}_2 \)

\[ \frac{75.0 \text{ g NH}_3 \left( \frac{1 \text{ mol}}{17.03 \text{ g}} \right)}{4.41 \text{ mol NH}_3} = 6.50 \text{ mol H}_2 \]

c) 46.5 g \( \text{H}_2 \)

\[ \frac{11.3 \text{ g H}_2}{2 \text{ mol H}_2} = 5.07 \text{ mol H}_2 \]

d) 11.3 g \( \text{H}_2 \)

\[ \frac{5.07 \text{ mol H}_2}{1 \text{ mol H}_2} = 13.3 \text{ g H}_2 \]

e) 13.3 g \( \text{H}_2 \)

12. Indium can react with oxygen gas to form indium(III) oxide. If this reaction has a percent yield of 63.8%, how many mol of indium must be reacted in excess oxygen to yield 3.00 mol of indium(III) oxide?

a) 5.74 mol \( \text{In} \)

b) 4.70 mol \( \text{In} \)

c) 9.40 mol \( \text{In} \)

d) 3.00 mol \( \text{In} \)

e) 6.00 mol \( \text{In} \)

\[ 2 \text{ In} \text{(s) + } \frac{3}{2} \text{O}_2 \text{(g) } \rightarrow \text{ In}_2\text{O}_3 \]

\[ \frac{6.00 \text{ mol In}}{0.638} = 9.40 \text{ mol In} \]
13. Consider the following reaction:

\[ 4 \text{C}_2\text{H}_3\text{OF} (g) + 9 \text{O}_2 (g) \rightarrow 8 \text{CO}_2 (g) + 6 \text{H}_2\text{O} (g) + 2 \text{F}_2 (g) \]

When 5.82 mol of C\textsubscript{2}H\textsubscript{3}OF and 10.35 mol of \text{O}_2 are initially present in the reaction mixture, how much of which reactant will remain after the reaction goes to completion?

a) 1.15 mol \text{O}_2 remain
b) 4.53 mol C\textsubscript{2}H\textsubscript{3}OF remain
c) 1.22 mol C\textsubscript{2}H\textsubscript{3}OF remain
d) 2.97 mol C\textsubscript{2}H\textsubscript{3}OF remain
e) 2.75 mol \text{O}_2 remain

\[ \text{mol} \text{C}_2\text{H}_3\text{OF} \times \frac{9 \text{ mol O}_2}{4 \text{ mol C}_2\text{H}_3\text{OF}} = 13.09 \text{ mol O}_2 \text{ would be used} \]

\[ \text{mol} \text{O}_2 \times \frac{4 \text{ mol C}_2\text{H}_3\text{OF}}{9 \text{ mol O}_2} = 4.60 \text{ mol C}_2\text{H}_3\text{OF} \text{ will be used} \]

\[ 5.82 \text{ mol} - 4.60 \text{ mol} = 1.22 \text{ mol} \text{ C}_2\text{H}_3\text{OF remain} \]

14. What mass of stones (C\textsubscript{s} = 0.841 J/g\cdot{°C}) that are heated to 815.0 °C must be added to 750.0 g of water (C\textsubscript{w} = 4.184 J/g\cdot{°C}) initially at 22.3 °C to raise the temperature of the water to the boiling point (100.0 °C)?

a) 793 g stones
b) 405 g stones
c) 445 g stones
(d) 366 g stones
e) 557 g stones

\[ m \times C_{s,\text{石头}} \times (T_\text{终} - T_\text{初}) = m \times C_{w,\text{水}} \times (T_\text{沸} - T_\text{初}) \]

\[ m \times 0.841 \frac{\text{J}}{\text{g} \cdot \text{°C}} \times (815.0 - 22.3) = m \times 4.184 \frac{\text{J}}{\text{g} \cdot \text{°C}} \times (100.0 - 22.3) \]
15. Consider the combustion of ethanol, C$_2$H$_5$OH:

\[ \text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 3\text{H}_2\text{O}(l) + 2\text{CO}_2(g) \]

Use Hess' law to calculate $\Delta H^\circ$ for this reaction. Equations that may be of use:

I. \[ \text{C}_2\text{H}_6(g) + 3\text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 2\text{CO}_2(g) \quad \Delta H^\circ = -1411 \text{ kJ} \]

II. \[ \text{C}_{\text{graphite}}(s) + 3\text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{C}_2\text{H}_5\text{OH}(l) \quad \Delta H^\circ \text{ not use} \]

III. \[ \text{C}_2\text{H}_6(g) + \text{H}_2\text{O}(l) \rightarrow \text{C}_2\text{H}_5\text{OH}(l) \quad \Delta H^\circ = -44 \text{ kJ} \]

a) 632 kJ

b) 44 kJ

c) \[ \text{C}_2\text{H}_5\text{OH} \rightarrow \text{C}_2\text{H}_4(g) + \text{H}_2\text{O} \quad \Delta H = +44 \text{ kJ} \]

d) \[ \text{C}_2\text{H}_4(g) + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{CO}_2 \quad \Delta H = -1411 \text{ kJ} \]

e) \[ \text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{CO}_2 \quad \Delta H = -1367 \text{ kJ} \]

16. Calculate the standard enthalpy for the following reaction:

\[ 4\text{NH}_3(g) + 7\text{O}_2(g) \rightarrow 4\text{NO}_2(g) + 6\text{H}_2\text{O}(l) \]

Given the following values:

$\Delta H^\circ$ for NH$_3(g) = -46 \text{ kJ/mol}$

$\Delta H^\circ$ for NO$_2(g) = 34 \text{ kJ/mol}$

$\Delta H^\circ$ for H$_2$O(l) = -286 kJ/mol

\[
\Delta H^\circ = \left[ \left( 4 \text{mol NH}_3 \times 34 \text{ kJ/mol} \right) + \left( 6 \text{mol H}_2\text{O} \times -286 \text{ kJ/mol} \right) \right] - \left[ \left( 4 \text{mol NH}_3 \times -46 \text{ kJ/mol} \right) + \left( 7 \text{mol O}_2 \times 0 \text{ kJ/mol} \right) \right]
\]

\[ = -1376 \text{ kJ} \]

d) -1396 kJ

e) Cannot be determined from the information provided
17. It takes 208.4 kJ of energy to remove 1 mol of electrons from an atom on the surface of rubidium (Rb) metal. What is the maximum wavelength of light capable of removing a single electron from an atom on the surface of solid Rb?

a) $9.532 \times 10^{-31}$ nm
b) $9.532 \times 10^{-11}$ nm
c) $9.532 \times 10^1$ nm
d) $5.740 \times 10^2$ nm
e) $9.740 \times 10^3$ nm

\[
\varepsilon = 208.4 \text{ kJ} \left( \frac{10^3 \text{ J}}{1 \text{ kJ}} \right) \left( \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ mol}} \right) = 3.416 \times 10^{-19} \text{ J/phot} \varepsilon \\
E = \frac{\hbar c}{\lambda} \quad \lambda = \frac{\hbar c}{E} = \frac{6.626 \times 10^{-34} \text{ J s}}{3.00 \times 10^8 \frac{\text{m}}{\text{s}}} = 3.416 \times 10^{-19} \frac{\text{m}}{\text{phot} \varepsilon} \\
= 5.74 \times 10^{-10} \text{ m} = 5.74 \text{ nm}
\]

18. The figure below can be labeled:

- a) a 2p$_x$ atomic orbital
- b) a 2p$_y$ atomic orbital
- c) a 5d atomic orbital
- d) a 5s atomic orbital
- e) a 4s atomic orbital
19. Give the complete electronic configuration for P.

a) \(1s^2 2s^2 2p^6 3s^1 3d^5\)

b) \(1s^2 1p^6 2s^2 2p^5\)

c) \(1s^2 2s^2 2p^6 3s^2 3p^3\)

d) \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3\)

e) \(1s^2 2s^2 2p^6 3s^2 3d^5\)

20. Choose the orbital diagram that represents the ground state of N.

a) \[
\begin{array}{ccc}
1s & 2s & 2p \\
\hline
- & 1 & 1 \\
\end{array}
\]

b) \[
\begin{array}{ccc}
1s & 2s & 2p \\
\hline
\frac{1}{2} & \frac{1}{2} & 1 \\
\end{array}
\]

c) \[
\begin{array}{ccc}
1s & 2s & 2p \\
\hline
1 & 1 & \frac{1}{2} \\
\end{array}
\]

d) \[
\begin{array}{ccc}
1s & 2s & 2p \\
\hline
\frac{1}{2} & \frac{1}{2} & 1 \\
\end{array}
\]

e) \[
\begin{array}{ccc}
1s & 2s & 2p \\
\hline
\frac{1}{2} & \frac{1}{2} & \frac{1}{2} \\
\end{array}
\]
21. Which of the following is **FALSE**?

a) Na is larger than Na⁺
b) Al is larger than Al³⁺
c) O is larger than O²⁻
d) O is larger than Ne
e) Xe is larger than Ne

22. Place the following in order of increasing radius.

\[ \text{Br}^- \quad \text{K}^+ \quad \text{Rb}^+ \]

a) Br⁻ < Rb⁺ < K⁺
b) K⁺ < Rb⁺ < Br⁻
c) Rb⁺ < Br⁻ < K⁺
d) Br⁻ < K⁺ < Rb⁺
e) Rb⁺ < K⁺ < Br⁻

23. Which of the following is FALSE?

a) Elements in the same group have the same valence electron configuration.  
   \[ \checkmark \]
b) Atomic radius of elements in a group increases as you go from top to bottom.  
   \[ \checkmark \]
c) Completely empty and completely full shells are very stable.  
   \[ \checkmark \]
d) First ionization energy of elements in the same group increases as you go from top to bottom.  
   \[ \checkmark \]
e) All of the above.
24. Chemical bonds form because:
   a) Doing so lowers the potential energy between the charged particles that compose atoms
   b) Doing so allows the bonded atoms to fit into a smaller volume
   c) Doing so raises the potential energy of the atoms that make up the molecule, which can subsequently release this energy by breaking the bond(s)
   d) Doing so allows the bonded atoms to have a greater density
   e) Doing so allows protons of one atom to combine with electrons of another to form neutrons

25. Under the Lewis model, which best explains why helium (He) obeys the duet rule instead of the octet rule?
   a) Helium is a noble gas
   b) Helium atoms almost never form covalent bonds
   c) Helium atoms are much less massive than most other elements
   d) The $n = 1$ quantum level fills with only two electrons
   e) None of the above statements are correct

26. All the carbon-carbon bonds in benzene (a Lewis structure is shown below) are known to be the same length. The reason for this is best supported by:

   ![Benzene Lewis structure]

   a) Hund's rule
   b) The concept of resonance
   c) The Aufbau principle
   d) The shape of carbon's 2s orbital
   e) The shape of carbon's 2p orbital
27. The Lewis Dot Structure of the hydroxide ion (OH⁻) depicts:

(A) There are no lone pairs of electrons.
(B) There is one lone pair of electrons.
(C) There are two lone pairs of electrons.
(D) There are three lone pairs of electrons.
(E) There are four lone pairs of electrons.

28. Choose the best Lewis structure for NH₄⁺.

(A) \[
\begin{array}{c}
H \\
\hline
N \\
H
\end{array}
\] 

(B) \[
\begin{array}{c}
H \\
\hline
N \\
H \\
H
\end{array}
\] 

(C) \[
\begin{array}{c}
H \\
\hline
N \\
H
\end{array}
\] 

(D) \[
\begin{array}{c}
H \\
\hline
N \\
H
\end{array}
\] 

(E) \[
\begin{array}{c}
H \\
\hline
N \\
H
\end{array}
\]
29. Choose the best Lewis structure for CH₂Cl₂.

A) \[ \text{H} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{H}} \]

B) \[ \begin{array}{c}
\overset{\cdot}{\text{C}} \\
\text{H} \\
\text{H}
\end{array} \]

C) \[ \text{H} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{C}} - \overset{\cdot}{\text{H}} \]

D) \[ \begin{array}{c}
\overset{\cdot}{\text{C}} \\
\text{H}
\end{array} \]

E) \[ \begin{array}{c}
\overset{\cdot}{\text{C}} = \overset{\cdot}{\text{C}} \\
\text{H}
\end{array} \]

30. The Lewis Dot Structure of H₂O depicts:

(A) There are no lone pairs of electrons.

(B) There is one lone pair of electrons.

(C) There are two lone pairs of electrons.

(D) There are three lone pairs of electrons.

(E) There are four lone pairs of electrons.

[TURN OVER FOR THE FINAL TWO QUESTIONS]
31. A student suggests a Lewis Structure for sulfate (SO₄²⁻) to have two double bonds and two single bonds to the central sulfur. The formal charge on the sulfur is:

a) -4  
b) -2  
c) 0  
d) +2  
e) +4

32. Which of the following would not exhibit any resonance structures?

a) O₃  
b) CF₄  
c) NO₃⁻  
d) SO₄²⁻  
e) CO₃²⁻