1. Predict boiling point based on lattice energy
2. Shapes of molecules
3. Exceptions to octet

1. On the sketch of the periodic table below that shows only the main group (representative) elements, draw an arrow from the element with the largest radius to the smallest radius.

<table>
<thead>
<tr>
<th>H</th>
<th>He</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>Be</td>
</tr>
<tr>
<td>Na</td>
<td>Mg</td>
</tr>
<tr>
<td>K</td>
<td>Ca</td>
</tr>
<tr>
<td>Rb</td>
<td>Sr</td>
</tr>
<tr>
<td>Cs</td>
<td>Ba</td>
</tr>
</tbody>
</table>

2. On the sketch of the periodic table below that shows only the main group (representative) elements, draw an arrow from the element with the largest first ionization energy to the element with the smallest first ionization energy.
3. Which one of the following electron configurations represents an excited state?
   a) 1s² 2s¹ 2p³  
   b) 1s² 2s² 2p³  
   c) 1s² 2s² 2p⁵  
   d) 1s² 2s² 2p⁶ 3s²  
   e) all of the above are excited state electron configurations  
   f) none of the configurations represent an excited state

4. Calculate ΔH for the reaction 2 C(s) + O₂(g) ------> 2 CO(g)
   given the following chemical equations and their respective enthalpy changes:

   C(s) + O₂(g) --------> CO₂(g)          ΔH = -393.5 kJ
   CO(g) + ½ O₂(g) --------> CO₂(g)      ΔH = -283.0 kJ

5. What is energy in Joules of one photon of yellow light with a wavelength of 589 nm? For full credit, show all work, use units, and report your answer to the correct number of significant figures.

6. Identify whether the following bonds are polar covalent, nonpolar covalent, or ionic. For polar covalent bonds, indicate the direction of the dipole moment using formalism.
7. Consider the following set of quantum numbers: \(n = 3, \ l = 2, \ ml = -2, \ ms = + 1/2\)

A. What orbital does this set of quantum numbers correspond to?
   a) 2p  b) 3s  c) 3p  d) 3d  e) 3f

B. How many electrons in a single atom can have this set of quantum numbers?
   a) 1  b) 2  c) 3  d) 4  e) unlimited

C. Which quantum number indicates the orientation of the orbital?
   a) \(n\)  b) \(l\)  c) \(ml\)  d) \(ms\)  e) none of them

D. Write the quantum number for another electron in the same orbital.

\[
\begin{array}{cccc}
\text{bond} & \text{nonpolar covalent, polar covalent, or ionic?} \\
\hline
\text{Cl – As} & \\
\text{Na – F} & \\
\hline
\end{array}
\]

8. Which of the following is the electron configuration of the Fe\(^{2+}\) ion?
   a) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^6 \ 4s^2\)
   b) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^4 \ 4s^2\)
   c) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^6\)
   d) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^5 \ 4s^2\)
   e) none of the above
9. Which of the following ions is isoelectronic with Ne?
   a) O2–
   b) F–
   c) Na+
   d) Mg2+
   e) all are isoelectronic with Ne

10. Which of the following ions has the smallest diameter?
    a) O2– b) F– c) Na+ d) Mg2+ e) all have the same diameter

11. Write the condensed electron configuration of a ground-state vanadium atom.

12. Write the correct ordering of the first ionization energy from smallest to largest for sodium, magnesium, silicon, and aluminum: Na, Mg, Si, Al

13. Consider the plot of the ionization energies for multiple ionizations of an atom in the third period. What is the element?
    a) Na
    b) Mg
    c) Al
    d) Si
    e) P
15. Which of the following phase change is **exothermic**?

a) H₂O(s) → H₂O(l)

b) H₂O(s) → H₂O(g)

c) H₂O(g) → H₂O(l)

d) H₂O(l) → H₂O(g)

e) None of the above

16. Which energy level can hold a **maximum** of 18 electrons?

a) n=1

b) n=2

c) n=3

d) n=4
e) none of the above

17. Which of the following electron configuration is not possible?

a) $1s^22s^22p^63s^13p^4$

b) $1s^22s^22p^63s^23p^54s^23d^{10}$

c) $1s^22s^22p^63s^23p^4$

d) $1s^22s^22p^63s^33p^4$

e) $1s^12s^22p^63s^23p^4$

18. Which one of the following represents an impossible set of quantum numbers for an electron in an atom? (given in the order $n$, $l$, $m_l$, and $m_s$)

a) 3, 2, -2, 1/2

b) 4, 0, 0, 1/2

c) 3, 3, 3, 1/2

d) 5, 3, 0, 1/2

e) 5, 3, 2, 1/2

19. Given the following reactions,

$2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ $\Delta H=-790kJ$

$S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H=-297kJ$

What is $\Delta H^\circ$ for the reaction $2SO_2 (g) + O_2 (g) \rightarrow 2SO_3 (g)$?

a) 196kJ

b) −196kJ

c) 1087kJ
For the next question, consider the covalent compound arsine, AsH₃. Arsine has the following Lewis structure:

\[ \text{H} \quad \cdot \quad \text{As} \quad \cdot \quad \text{H} \]

20. The central atom in arsine has ___ bonding pair(s) of electrons and ___ lone pair(s) of electrons.

A. one, one  
B. one, three  
C. zero, one  
D. three, one  
E. three, two

21. Fill in the electron configuration for Fluorine.
22. A student must use 225 mL of hot water in a lab procedure. Calculate the amount of heat required to raise the temperature of 225 mL of water from 20.0 °C to 100.0 °C. (Density of water is 1.00 g/mL)
### TABLE 4.1 Solubility Guidelines for Common Ionic Compounds in Water

<table>
<thead>
<tr>
<th>Soluble Ionic Compounds</th>
<th>Important Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing</td>
<td></td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>None</td>
</tr>
<tr>
<td>C₂H₃O₂⁻</td>
<td>None</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>Compounds of Ag⁺, Hg₂²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>Br⁻</td>
<td>Compounds of Ag⁺, Hg₂²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>I⁻</td>
<td>Compounds of Ag⁺, Hg₂²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>Compounds of Sr²⁺, Ba²⁺, Hg₂²⁺, and Pb²⁺</td>
</tr>
</tbody>
</table>

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<tr>
<th>Insoluble Ionic Compounds</th>
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<tbody>
<tr>
<td>Compounds containing</td>
<td></td>
</tr>
<tr>
<td>S²⁻</td>
<td>Compounds of NH₄⁺, the alkali metal cations, and Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>Compounds of NH₄⁺ and the alkali metal cations</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>Compounds of NH₄⁺ and the alkali metal cations</td>
</tr>
<tr>
<td>OH⁻</td>
<td>Compounds of the alkali metal cations, and NH₄⁺, Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
</tbody>
</table>

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density of water: 1.00 g/mL
1 cal = 4.184 joules
\[ c = 3.00 \times 10^{10} \text{ cm/s} \text{ or } 3.00 \times 10^{8} \text{ m/s} \]
h is 6.63 \times 10^{-34} \text{ joule sec}
\[ \frac{1}{\lambda} = R_{\mu} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) n_2 > n_1 \]
\[ R_{\mu} = 1.096776 \times 10^7 \text{ m}^{-1} \]
\[ \mu = Q \tau \]
\[ 3.34 \times 10^{30} \text{ C-m} = 1 \text{ Debye} \]
electronic charge: \[ e = 1.60 \times 10^{-19} \text{ C} \]
\[ E_{\text{in}} = -2.18 \times 10^{-18} \text{ J} \left( \frac{1}{\pi^2} \right) \]
\[ \Delta x \Delta mv \geq \frac{h}{4 \pi} \]