

QUESTIONS ON THE PROPERTIES OF LIGHT

$$1. \quad c = \lambda \nu, \text{ or } \lambda = \frac{c}{\nu} \quad \text{where } c = 2.9979 \times 10^8 \text{ m/s}$$

$$a) \quad \lambda = \frac{2.9979 \times 10^8 \text{ m/s}}{5.09 \times 10^{14} \text{ s}^{-1}} = 5.89 \times 10^{-7} \text{ m}$$

$$b) \quad \lambda = \frac{2.9979 \times 10^8 \text{ m/s}}{8.6 \times 10^{12} \text{ s}^{-1}} = 3.5 \times 10^{-5} \text{ m}$$

$$c) \quad \lambda = \frac{2.9979 \times 10^8 \text{ m/s}}{2.0 \times 10^9 \text{ s}^{-1}} = 1.5 \times 10^{-1} \text{ m}$$

$$2. \quad c = \lambda \nu, \text{ or } \nu = \frac{c}{\lambda} \quad \text{and} \quad E = h \nu \quad \text{or} \quad E = \frac{h \cdot c}{\lambda} \quad \text{where } h = 6.6262 \times 10^{-34} \text{ J}\cdot\text{s}$$

$$a) \quad \nu = \frac{2.9979 \times 10^8 \text{ m/s}}{1.8 \times 10^{-2} \text{ m}} = 1.7 \times 10^{10} \text{ s}^{-1}$$

$$E = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{1.8 \times 10^{-2} \text{ m}} = 1.1 \times 10^{-23} \text{ J}$$

$$b) \quad \nu = \frac{2.9979 \times 10^8 \text{ m/s}}{4.80 \times 10^{-8} \text{ m}} = 6.25 \times 10^{15} \text{ s}^{-1}$$

$$E = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{4.80 \times 10^{-8} \text{ m}} = 4.14 \times 10^{-18} \text{ J}$$

$$c) \quad \nu = \frac{2.9979 \times 10^8 \text{ m/s}}{3.05 \times 10^{-7} \text{ m}} = 9.83 \times 10^{14} \text{ s}^{-1}$$

$$E = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{3.05 \times 10^{-7} \text{ m}} = 6.51 \times 10^{-19} \text{ J}$$

$$3. \quad E = h \nu, \text{ or } \nu = \frac{E}{h} \quad \text{and} \quad E = \frac{h \cdot c}{\lambda} \quad \text{or} \quad \lambda = \frac{h \cdot c}{E}$$

$$a) \quad \nu = \frac{6.95 \times 10^{-20} \text{ J}}{6.6262 \times 10^{-34} \text{ J}\cdot\text{s}} = 1.05 \times 10^{14} \text{ s}^{-1}$$

and,

$$\lambda = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{6.95 \times 10^{-20} \text{ J}} = 2.86 \times 10^{-6} \text{ m}$$

$$b) \nu = \frac{5.96 \times 10^{-26} \text{ J}}{6.6262 \times 10^{-34} \text{ J}\cdot\text{s}} = 8.99 \times 10^7 \text{ s}^{-1} \quad \text{and,}$$

$$\lambda = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{5.96 \times 10^{-26} \text{ J}} = 3.33 \text{ m}$$

$$c) \nu = \frac{2.00 \times 10^{-15} \text{ J}}{6.6262 \times 10^{-34} \text{ J}\cdot\text{s}} = 3.02 \times 10^{18} \text{ s}^{-1} \quad \text{and,}$$

$$\lambda = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{2.00 \times 10^{-15} \text{ J}} = 9.93 \times 10^{-11} \text{ m}$$

$$4. a) \bar{\nu} = \frac{1}{\lambda} ; \text{ therefore...}$$

$$\bar{\nu} = \frac{1}{150 \times 10^{-12} \text{ m}} \times \frac{1 \text{ m}}{10^2 \text{ cm}} = 6.67 \times 10^7 \text{ cm}^{-1} \quad \text{This is in the X-ray region.}$$

$$b) E = h \cdot c \cdot \bar{\nu} ; \text{ therefore...}$$

$$E = 6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s} \times \frac{4000}{10^{-2} \text{ m}} = 7.946 \times 10^{-20} \text{ J} \quad \text{This is in the IR region.}$$

$$c) \lambda = \frac{c}{\nu} ; \text{ therefore...}$$

$$\lambda = \frac{2.9979 \times 10^8 \text{ m/s}}{1.50 \times 10^{15} \text{ s}^{-1}} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 2.00 \times 10^2 \text{ nm} \quad \text{This is in the UV region.}$$

$$d) E = \frac{h \cdot c}{\lambda} ; \text{ therefore...}$$

$$E = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{10 \times 10^{-2} \text{ m}} = 2.0 \times 10^{-24} \text{ J} \quad \text{This is in the MW region.}$$

$$e) \nu = \frac{E}{h} ; \text{ therefore...}$$

$$\nu = \frac{1.0 \times 10^{-18} \text{ J}}{6.6262 \times 10^{-34} \text{ J}\cdot\text{s}} = 1.5 \times 10^{15} \text{ s}^{-1} \quad \text{This is in the UV region.}$$

$$f) \bar{\nu} = \frac{\nu}{c} ; \text{ therefore...}$$

$$\bar{\nu} = \frac{1.20 \times 10^{14} \text{ s}^{-1}}{2.9979 \times 10^8 \text{ m/s}} \times \frac{1 \text{ m}}{10^2 \text{ cm}} = 4.00 \times 10^3 \text{ cm}^{-1} \quad \text{This is in the IR region.}$$

5. $E = h \cdot \nu = \frac{h \cdot c}{\lambda}$; therefore...

$$E = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{650 \times 10^{-9} \text{ m}} = 3.05_6 \times 10^{-19} \text{ J or } 3.05_6 \times 10^{-19} \text{ J/photon}$$

$$E_{\text{tot}} = \frac{E}{\text{photon}} \times \frac{\# \text{photons}}{\text{mole}} = \frac{3.05_6 \times 10^{-19} \text{ J}}{\text{photon}} \times \frac{6.022 \times 10^{23} \text{ photons}}{\text{mol}} = 1.84 \times 10^5 \text{ J/mol}$$

6. X-rays: $E = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{0.15 \times 10^{-9} \text{ m}} = 1.3_2 \times 10^{-15} \text{ J}$

Microwaves: $E = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{1.00 \times 10^{-3} \text{ m}} = 1.98_6 \times 10^{-22} \text{ J}$

$$\frac{\text{X-ray Energy}}{\text{Microwave Energy}} = \frac{1.3_2 \times 10^{-15} \text{ J}}{1.98_6 \times 10^{-22} \text{ J}} = 6.6 \times 10^6 !$$

7. $E_{\text{photon}} = h \cdot \nu = \frac{h \cdot c}{\lambda}$; therefore...

$$E = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{500 \times 10^{-9} \text{ m}} = 3.97_3 \times 10^{-19} \text{ J}$$

$$\text{Number of photons} = 1.59 \times 10^{-17} \text{ J} \times \frac{1 \text{ photon}}{3.97_3 \times 10^{-19} \text{ J}} = 40 \text{ photons (an integer)}$$

8. $E_{\text{absorbed}} = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{460 \times 10^{-9} \text{ m}} = 4.31_8 \times 10^{-19} \text{ J}$

$$E_{\text{emitted}} = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{660 \times 10^{-9} \text{ m}} = 3.01_0 \times 10^{-19} \text{ J}$$

$$\Delta E = E_{\text{absorbed}} - E_{\text{emitted}} = 4.31_8 \times 10^{-19} \text{ J} - 3.01_0 \times 10^{-19} \text{ J} = 1.31 \times 10^{-19} \text{ J}$$

9. a) $E = h \cdot \nu = (6.6262 \times 10^{-34} \text{ J}\cdot\text{s}) \times (1.3 \times 10^{15} \text{ s}^{-1}) = 8.6 \times 10^{-19} \text{ J}$

b) UV (~100nm): $E = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{100 \times 10^{-9} \text{ m}} = 1.99 \times 10^{-18} \text{ J}$

Photoelectric effect is observed: Energy of UV photon exceeds threshold energy for Pt.

b) IR (~1.00 × 10⁻³ cm): $E = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{1.00 \times 10^{-5} \text{ m}} = 1.99 \times 10^{-20} \text{ J}$

Photoelectric effect is not observed: Energy of IR photon is less than threshold energy for Pt.

- c) Increasing the intensity increases the number of photons and not the energy per photon. Therefore, an increase in the UV intensity would increase the photoelectric effect ($\therefore$ current increases), while an increase in IR intensity would have no effect (more photons, but each with insufficient energy to remove an e^- from Pt).
- d) If the λ of UV light decrease, the energy content per photon increases. The energy in excess of the threshold energy for the removal of an e^- would be added to the kinetic energy of the liberated e^- .

10.
$$\frac{\text{Threshold Energy of Cs}}{E} = h \cdot \nu = (6.6262 \times 10^{-34} \text{ J}\cdot\text{s}) \times (4.60 \times 10^{14} \text{ s}^{-1}) = 3.05 \times 10^{-19} \text{ J}$$

$$E_{\text{photon}} (\lambda = 540 \text{ nm}) = \frac{h \cdot c}{\lambda} = \frac{(6.6262 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m/s})}{540 \times 10^{-9} \text{ m}} = 3.68 \times 10^{-19} \text{ J}$$

$$E_{\text{photon}} = E_b \text{ (threshold energy)} + KE(e^-)$$

$$\text{or } KE(e^-) = E_{\text{photon}} - E_b = 3.68 \times 10^{-19} \text{ J} - 3.05 \times 10^{-19} \text{ J} = 6.3 \times 10^{-20} \text{ J}$$

11. Electrons in energy levels close to the nucleus are of the lower energy and have lower principal quantum number, n . As the value of n increases, the electrons in these levels are further from the nucleus and are of higher energy. Therefore, energy is absorbed by the electron as it is excited from a lower energy level to a higher energy level and energy is released in the reverse process.

a) $n=2$ to $n=1$ **Energy is released** b) $n=2$ to $n=4$ **Energy is absorbed**

c) $n=2$ to $n=\infty$ **Energy is absorbed**

12. a)
$$\Delta E_{i \rightarrow f} = 2.18 \times 10^{-18} \text{ J} \left[\frac{1}{n_f^2} - \frac{1}{n_i^2} \right]$$

$$\Delta E_{4 \rightarrow 3} = 2.18 \times 10^{-18} \text{ J} \left[\frac{1}{3^2} - \frac{1}{4^2} \right] = 1.06 \times 10^{-19} \text{ J}$$

$$\Delta E_{3 \rightarrow 2} = 2.18 \times 10^{-18} \text{ J} \left[\frac{1}{2^2} - \frac{1}{3^2} \right] = 3.03 \times 10^{-19} \text{ J}$$

$$\Delta E_{2 \rightarrow 1} = 2.18 \times 10^{-18} \text{ J} \left[\frac{1}{1^2} - \frac{1}{2^2} \right] = 1.64 \times 10^{-18} \text{ J}$$

Energy differences between adjacent levels DECREASE in magnitude as n INCREASES.
(Energy levels are progressively closer together as n increases.)

$$\text{b) } \Delta E = \frac{h \cdot c}{\lambda} \quad \text{or} \quad \lambda = \frac{h \cdot c}{\Delta E} \quad \text{Spectral Region}$$

$$\lambda_{4-3} = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{1.06 \times 10^{-19} \text{ J}} = 1.87 \times 10^{-6} \text{ m} \quad \text{UV}$$

$$\lambda_{3-2} = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{3.03 \times 10^{-19} \text{ J}} = 6.56 \times 10^{-7} \text{ m} \quad \text{Visible}$$

$$\lambda_{2-1} = \frac{6.6262 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.9979 \times 10^8 \text{ m/s}}{1.64 \times 10^{-18} \text{ J}} = 1.21 \times 10^{-7} \text{ m} \quad \text{UV}$$



$$\Delta E_{1-\infty} = 2.18 \times 10^{-18} \text{ J} \left[\frac{1}{1^2} - \frac{1}{(\infty)^2} \right] = 2.18 \times 10^{-18} \text{ J}$$

b) Li^{+2} is a one electron system with 3 protons in the nucleus. The hydrogen atom is a one electron system with only one proton in the nucleus. The force of attraction of the three protons in Li^{+2} for an electron is obviously greater than the attraction of the lone proton in the H-atom for an electron. The ionization energy is much larger for Li^{+2} compared to the H-atom. In fact the ionization energy for one electron systems is proportional to the square of the nuclear charge (Z^2) and therefore the ionization potential of Li^{+2} is nine times that of the H-atom.

14. a) **For particles** $\lambda = \frac{h}{m \cdot v} = \frac{6.6262 \times 10^{-34} \text{ kg}\cdot\text{m}^2/\text{s}}{1.00 \times 10^{-8} \text{ kg} \times 10.0 \text{ m/s}} = 6.63 \times 10^{-27} \text{ m}$

This wavelength is far too small to be detected by any device existing at this time. Larger objects have even larger masses and therefore, have still smaller wavelengths.

b) $v_e = 0.100 \times 2.9979 \times 10^8 \text{ m/s} = 3.00 \times 10^7 \text{ m/s}$

$$\lambda = \frac{h}{m \cdot v} = \frac{6.6262 \times 10^{-34} \frac{\text{kg}\cdot\text{m}^2}{\text{s}}}{9.11 \times 10^{-31} \text{ kg} \times 3.00 \times 10^7 \text{ m/s}} = 2.43 \times 10^{-11} \text{ m}$$

14. DIFFERENCES

- a. Bohr Theory: The electron may be found only in one of a fixed set of closed circular orbits.
 Wave Mechanics: The electron may be found anywhere outside the nucleus, but, there are regions called orbitals which have the highest probability of containing the electron.
- b. Bohr Theory: The electron orbit is a planar figure (i.e., a circle).
 Wave Mechanics: The orbitals represent to 3-D regions (e.g., spherical, dumbbell-shaped, etc).
- c. Bohr Theory: The position and velocity of the electron can be described with certainty.
 Wave Mechanics: Certainty is replaced by probability in describing the position and velocity.

SIMILARITIES

Electron energies are quantized.

Radii of Bohr orbits correspond to distances from the nucleus where the probabilities of finding the electron are high .

QUESTIONS ON ELECTRON CONFIGURATIONS

16. Recall that the quantum numbers have allowed values of:

$$\begin{aligned} n &= 1, 2, 3, 4, \dots \\ \ell &= 0, 1, 2, 3, \dots (n-1) \\ m_{\ell} &= -\ell, (-\ell + 1), \dots 0 \dots (\ell-1), \ell \\ s &= +\frac{1}{2}, -\frac{1}{2} \end{aligned}$$

- a) $n = 4$
 $\ell = 0$
 $m_{\ell} = 0$ only possible value
 $s = +\frac{1}{2}, -\frac{1}{2}$ } ($\therefore$ 2 sets of 4 quantum numbers $\therefore$ 2 e $^{-}$)
- b) $n = 4$
 $\ell = 1$
 $m_{\ell} = \left\{ \begin{array}{c} -1 \\ 0 \\ +1 \end{array} \right\}$
 $s = \pm \frac{1}{2}$ } ($\therefore$ 6 sets of 4 quantum numbers $\therefore$ 6 e $^{-}$)
- c) $n = 4$
 $\ell = 2$
 $m_{\ell} = \left\{ \begin{array}{c} -2 \\ -1 \\ 0 \\ +1 \\ +2 \end{array} \right\}$
 $s = \pm \frac{1}{2}$ } ($\therefore$ 10 sets of 4 quantum numbers $\therefore$ 10 e $^{-}$)
- d) $n = 4$
 $\ell = 3$
 $m_{\ell} = \left\{ \begin{array}{c} -3 \\ -2 \\ -1 \\ 0 \\ +1 \\ +2 \\ +3 \end{array} \right\}$
 $s = \pm \frac{1}{2}$ } ($\therefore$ 14 sets of 4 quantum numbers $\therefore$ 14 e $^{-}$)
- e) $n = 4$
 $\ell = 3$
 $m_{\ell} = 0$
 $s = \pm \frac{1}{2}$ } ($\therefore$ 2 sets of 4 quantum numbers $\therefore$ 2 e $^{-}$)

17. Recall that the quantum numbers have allowed values of:

$$n = 1, 2, 3, 4, \dots$$

$$\ell = 0, 1, 2, 3, \dots (n-1)$$

$$m_\ell = -\ell, (-\ell + 1), \dots, 0, 1, \dots, \ell$$

$$s = +\frac{1}{2}, -\frac{1}{2}$$

Note that ℓ is defined in terms of n , and that m_ℓ is defined in terms of ℓ and therefore m_ℓ is also limited by the value of n .

a) $n = 2, \ell = 1, m_\ell = 0$

Fits the above requirements: $\therefore$ allowable

b) $n = 2, \ell = -1$

The minimum value of ℓ is 0: $\therefore$ not allowable

c) $n = 3, \ell = 0, m_\ell = 0$

Allowable

d) $n = 3, \ell = 1, m_\ell = -1$

Allowable

e) $n = 2, \ell = 0, m_\ell = -1$

If $\ell = 0$, m_ℓ must equal 0: $\therefore$ not allowable

f) $n = 3, \ell = 3, m_\ell = 2$

ℓ cannot equal or exceed n : $\therefore$ not allowable

18. The ℓ quantum number designates the orbital type ($\ell = 0$ for s orbitals; $\ell = 1$ for p orbitals; $\ell = 2$ for d orbitals; $\ell = 3$ for f orbitals; etc.) and the n quantum number designates the principal electronic shell or energy level in which the orbital is found.

a) $n = 2, \ell = 1, m_\ell = -1$ 2p orbital b) $n = 4, \ell = 0, m_\ell = 0$ 4s orbital

c) $n = 5, \ell = 2, m_\ell = 0$ 5d orbital

19. The ℓ quantum number designates the orbital type ($\ell = 0$ for s orbitals; $\ell = 1$ for p orbitals; $\ell = 2$ for d orbitals; $\ell = 3$ for f orbitals; etc.) and the n quantum number designates the principal electronic shell or energy level in which the orbital is found.

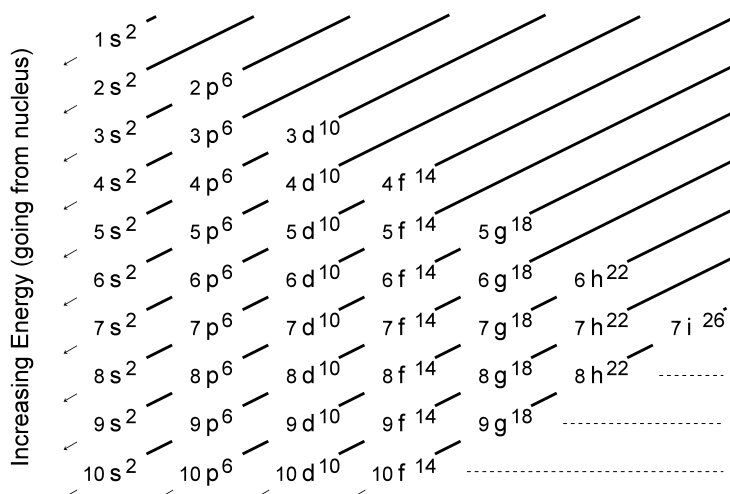
a) 3s, $n = 3, \ell = 0$

b) 4p, $n = 4, \ell = 1$

c) 5d, $n = 5, \ell = 2$

20. The Pauli exclusion principle states that no two electrons in a given atom may be assigned the same set of four quantum numbers.

21. AUFBAU ORDER



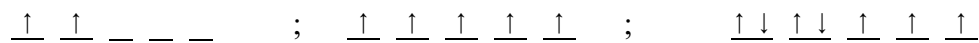
22. Hund's rule states that orbitals of a given sublevel are filled singly (with spins parallel) before pairing begins. (Since this results in the production of the maximum number of parallel spins, it is often referred to as maximum spin or maximum multiplicity.)
23. There are 3 orbitals in each p-sublevel, 5 orbitals in each d-sublevel, and 7 orbitals in each f-sublevel.

For 2, 4, and 5 electrons in a p-sublevel, the configurations would look like:



The 2 and 4 electron configurations are of equal paramagnetic character as each contains 2 unpaired electrons. The 5 electron configuration is of lower paramagnetic character as it contains only 1 unpaired electron.

For 2, 5, and 7 electrons in a d-sublevel, the configurations would look like:



The 2 electron configuration is of lowest paramagnetic character as it has only 2 unpaired electrons. The 7 electron configuration is of next highest paramagnetic character with 3 unpaired electrons and the 5 electron configuration with its 5 unpaired electrons is of highest paramagnetic character.

For 6, 7, and 8 electrons in an f-sublevel, the configurations would be:



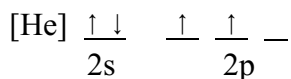
Here, the 6 and 8 electron configurations are of equal paramagnetic character with 6 unpaired electrons each. The 7 electron configuration is of higher paramagnetic character as it has one additional unpaired electron.

24. Probably the most direct approach is to convert each set of quantum numbers to an orbital designation and then arrange the orbitals in accordance with the aufbau order.

a) $n = 4, \ell = 0, m_\ell = 0, s = +1/2$	4s orbital	
b) $n = 3, \ell = 1, m_\ell = -1, s = -1/2$	3p orbital	
c) $n = 3, \ell = 2, m_\ell = 0, s = +1/2$	3d orbital	} degenerate orbitals
d) $n = 3, \ell = 2, m_\ell = -2, s = -1/2$	3d orbital	
e) $n = 3, \ell = 0, m_\ell = 0, s = -1/2$	3s orbital	

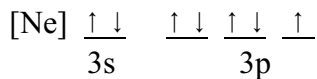
Order of increasing energy: $3s < 3p < 4s < 3d$

25. a) C: $1s^2 2s^2 2p^2$



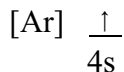
paramagnetic

b) Cl: $[\text{Ne}] 3s^2 3p^5$



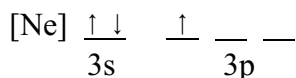
paramagnetic

c) K: $[\text{Ar}] 4s^1$



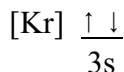
paramagnetic

d) Al: $[\text{Ne}] 3s^2 3p^1$



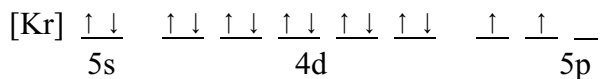
paramagnetic

e) Sr: $[\text{Kr}] 5s^2$



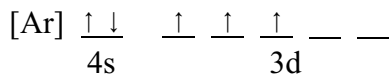
diamagnetic

f) Sn: $[\text{Kr}] 5s^2 4d^{10} 5p^2$



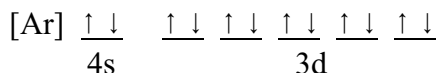
paramagnetic

g) V: $[\text{Ar}] 4s^2 3d^3$



paramagnetic

h) Zn: $[\text{Ar}] 4s^2 3d^{10}$



diamagnetic

26. Mo and Au are in the same families as Cr and Cu, respectively. These groups exhibit electron configurations which are exceptions to the configurations predicted from a strict aufbau order of filling. These systematic exceptions reflect the extra stability associated with half-filled and completely filled d-sublevels.

27. a) Na ($Z = 11$) $1s^2 2s^2 2p^6 3s^1$

b) P ($Z = 15$) $1s^2 2s^2 2p^6 3s^2 3p^3$

c) Te ($Z = 52$) $[\text{Kr}] 5s^2 4d^{10} 5p^4$

d) Zr ($Z = 40$) $[\text{Kr}] 5s^2 4d^2$

e) I ($Z = 53$) $[\text{Kr}] 5s^2 4d^{10} 5p^5$

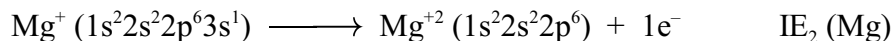
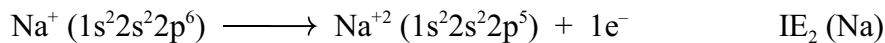
f) Bi ($z = 83$) $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^3$

28. a) $[\text{Ne}]3s^23p^3 = 15 e^- \therefore$ the element is P
 b) $[\text{Ar}]4s^23d^{10}4p^4 = 34 e^- \therefore$ the element is Se
 c) $[\text{Ar}]4s^13d^5 = 24 e^- \therefore$ the element is Cr
 d) $[\text{Kr}]5s^24d^2 = 40 e^- \therefore$ the element is Zr
 e) $[\text{Xe}]6s^1 = 55 e^- \therefore$ the element is Cs
 f) $[\text{Xe}]6s^24f^{14}5d^{10}6p^6 = 86 e^- \therefore$ the element is Rn
29. a) ns^2 - Group II A
 b) ns^2np^2 - Group IV A
 c) ns^2np^6 - Group VIII A
 d) $ns^2(n-1)d^2$ - Group IV B (Ti family)
 e) $ns^2(n-1)d^5$ - Group VII B (Mn Family)
30. a) $1s^22s^12p^1$ One 2s electron has been promoted to an empty 2p orbital.
 b) $[\text{Ne}]3s^23p^23d^2$ Two 3p electrons have been promoted to the 3d orbitals.
 c) $[\text{Ar}]4s^13d^{10}4p^3$ One 4s electron has been promoted to a 4p orbital.

QUESTIONS ABOUT PERIODIC TRENDS

31. Recall the atomic radii trends:
- | | | | |
|---|---|----------------------------|---|
| - Decrease across a period from L $\longrightarrow$ R | } | a) $\text{Sr} > \text{Sb}$ | b) $\text{Cl} < \text{I}$ |
| - Increase down a family or group | | c) $\text{P} < \text{Ge}$ | d) $\text{Si} ? \text{Se}$ Trends conflict $\therefore$ can't tell. |
32. a) $\text{Ba} > \text{Au} > \text{Ag} > \text{Te} > \text{Ar} > \text{H}$ b) $\text{Rb} > \text{K} > \text{Ca} > \text{Ti} > \text{Br} > \text{Cl}$
33. Removal of the first electron results in a decrease in size due to decrease in the number of electron to electron repulsions. Consequently, the remaining electrons are held more tightly because they are closer to the nucleus than before. Therefore, the removal of a second electron requires more energy.
34. For the first electron,
- $$\text{Na} (1s^22s^22p^63s^1) \longrightarrow \text{Na}^+ (1s^22s^22p^6) + 1e^- \quad \text{IE}_1 (\text{Na})$$
- $$\text{Mg} (1s^22s^22p^63s^2) \longrightarrow \text{Mg}^+ (1s^22s^22p^63s^1) + 1e^- \quad \text{IE}_1 (\text{Mg})$$
- The first electron lost by a Na atom is a 3s electron. The same is true for a Mg atom. However, because the effective nuclear charge (Z^*) is greater for the valence electrons of a Mg atom than for the valence electron of a Na atom, the 3s electron of Na is more easily removed than one of the 3s electrons of Mg, i.e., $\text{IE}_1 (\text{Na}) < \text{IE}_1 (\text{Mg})$.

For the second electron,



The second electron lost by a Mg atom is also a 3s electron, whereas, the second electron lost by Na must now come from an orbital which is on average closer to the nucleus and thus a more energetically stable orbital. As a result $\text{IE}_2 (\text{Na}) > \text{IE}_2 (\text{Mg})$.

35. Nonmetals generally have a greater number of valence electrons and these electrons typically experience a greater effective nuclear charge (Z^*) than the valence electrons of metals. Consequently, IE_1 of metals is generally less than IE_1 of nonmetals.

36. Recall the general trends for the first ionization energies:
- | | | | |
|--|---|----------------------------|----------------------------|
| -Increases across a period ($L \longrightarrow R$) | } | a) $\text{Ar} > \text{K}$ | b) $\text{S} > \text{As}$ |
| -Increases up a family or group | | c) $\text{Cs} < \text{Ba}$ | d) $\text{Cl} > \text{I}$ |
| | | e) $\text{S} < \text{P}$ | f) $\text{Sr} < \text{Sn}$ |

Note: Part (e) is an exception to the general trend. The electron configuration of P is $[\text{Ne}]3s^2 3p^3$. The half-filled 3p sublevel achieves some added stability as a consequence of this symmetric arrangement. Therefore, the first ionization energy of P is higher than normally expected. In addition, the loss of an electron from the valence shell of S ($[\text{Ne}]3s^2 3p^4$) would leave the 3p sublevel half-filled and therefore with some additional stability. Consequently the first ionization energy of S is lower than would normally be expected.

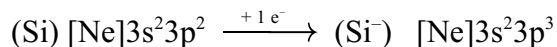
37. a) $\text{Cs} < \text{Rb} < \text{Ca} < \text{Sc} < \text{Se} < \text{Br} < \text{O} < \text{F} < \text{He}$

- b) Metallic character may be viewed as the tendency of an element to give up valence electrons, i.e., the inverse of ionization energy.

Decreasing metallic character: $\text{Cs} > \text{K} > \text{Ca} > \text{Al} > \text{Te} > \text{I} > \text{O} > \text{F}$

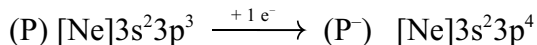
38. a) Sulfur and chlorine are both in the same period with six and seven valence electrons, respectively. The effective nuclear charge experienced by an added electron would be greater for chlorine than for the added electron in sulfur. Consequently, the electron affinity of chlorine should be higher than that of sulfur. Note, that the addition of an electron to chlorine would yield an electron configuration with a filled sublevel (i.e., $[\text{Ne}]3s^2 3p^6$). This would serve to further increase the electron affinity of chlorine over that of sulfur.

- b) This is another "exception" to the general trend in electron affinities. The addition of one electron to a silicon atom would yield an electron configuration with a half-filled 3p sublevel:



thus making the electron affinity of Si higher than normally expected.

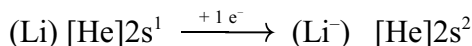
The addition of one electron to a phosphorus atom would disrupt a half-filled 3p sublevel:



thus making the electron affinity of P lower than normally expected. These two effects combine to cause the electron affinity of phosphorus to be lower than that of silicon (contrary to the general trend).

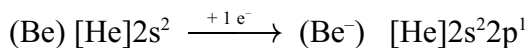
- c) This is a similar situation to part (b).

The addition of an electron to a Li atom would complete the 2s sublevel:



thus making the electron affinity of Li higher than normally expected. (Continued)

The addition of an electron to a Be atom would require placement of the added electron into the vacant 2p sublevel. This would disrupt the configuration which has all sublevels completely filled:



thus making the electron affinity of Be lower than normally expected. These two effects combine to cause the electron affinity of Be to be lower than that of Li (contrary to the general trend).

39. Increasingly negative ΔH_{EA} :

Ba, Cs (Ba has filled sublevel) , P, Si (P has half filled sublevel!) , Cl , F

40. Recall the electronegativity trends: EN increases across period (L $\rightarrow$ R)

EN increases up a group or family

- a) $\text{C} < \text{N} < \text{O}$ b) $\text{K} < \text{Fe} < \text{Cu}$ c) $\text{Br} < \text{Cl} < \text{F}$ d) $\text{Si} < \text{C} < \text{N}$ e) $\text{Cs} < \text{Ga} < \text{B}$

41. a) Largest atomic radius; Na b) Highest first ionization energy; Ar
c) Most reactive metal; Na d) Most reactive nonmetal; Cl
e) Least reactive; Ar

QUESTIONS ABOUT IONIC SUBSTANCES

42. a) Formation of a positive ion requires the removal of one or more electrons from the neutral parent atom. The remaining electrons experience a decrease in the degree of electron-electron repulsion. Consequently, the attractions between the nucleus and the valence electrons now exceed the repulsions between the valence electrons which cause the valence electrons to be drawn inward toward the nucleus and thus decreasing the radius of the ion relative to the parent atom.
- b) Formation of a negative ion requires the addition of one or more electrons to the neutral parent atom. The additional electron(s) cause an increase in the degree of electron-electron repulsions. Consequently, the repulsions between the valence electrons now exceed the attractions between the nucleus and the valence electrons. This causes the valence electrons to be pushed outward away from the nucleus and thus increasing the radius of the ion relative to the parent atom.

43. We should expect the hydrogen ion, H^+ , to be the smallest of all chemical species, since it consists of a lone proton. The hydride ion, H^- , should be larger than the He atom. Although He and H^- both have two electrons ($1s^2$), the H^- has only one proton to interact with the two electrons while He has two protons.
44. Recall that the elements in the s and p blocks tend to form ions with the noble gas configurations.
- a) Mg is in group IIA. These elements have relatively low ionization energies and unfavorable electron affinities and thus form cations. Mg can achieve a noble gas configuration (NGC) by losing two electrons. Therefore, Mg is expected to form a +2 ion (Mg^{+2}).
 - b) K is in group IA. Like Mg, K is expected to form a cation but it will achieve an NGC by loss of one electron and therefore K is expected to form a +1 ion (K^+).
 - c) Al is in group IIIA. Like Mg and K, Al is expected to form a cation but must lose three electrons to achieve NGC. Therefore Al is expected to form Al^{+3} ions.
 - d) P is in group VA. These elements have relatively high ionization energies and favorable electron affinities and therefore form anions. P can achieve an NGC by gaining three electrons. Therefore, P is expected to form an anion with a -3 charge (P^{-3}).
 - e) I is in group VIIA. Like P, I is expected to form an anion but will achieve an NGC by gaining one electron and thus I is expected to form a -1 ion (I^-).
 - f) Te is in group VIA. Like P and I, Te is expected to form an anion but with a -2 charge (Te^{-2}).
45. Ionic compounds form only when electrical neutrality is maintained. Therefore the cations and anions must combine in such number that the compound is electrically neutral (i.e., carry no charge).
- a) Ca is expected to form Ca^{+2} and S is expected to form S^{-2} . Therefore, the compound expected from Ca and S_8 is CaS .
 - b) Li is expected to form Li^+ and P is expected to form P^{-3} . Therefore, the compound expected from Li and P_4 is Li_3P .
 - c) Ga is expected to form both Ga^+ and Ga^{+3} while F is expected to form F^- . Therefore, the compounds expected from Ga and F_2 are GaF and GaF_3 .
 - d) Mg is expected to form Mg^{+2} and N is expected to form N^{-3} . Therefore, the compound expected from Mg and N_2 is Mg_3N_2 .
 - e) Al is expected to form Al^{+3} and O is expected to form O^{-2} . Therefore, the compound expected from Al and O_2 is Al_2O_3 .
 - f) Na is expected to form Na^+ and H is expected to form H^- . Therefore, the compound expected from Na and H_2 is NaH .

46. a) $\text{O} ([\text{He}]2s^22p^4) \longrightarrow \text{O}^{2-} ([\text{He}]2s^22p^6 \text{ or } [\text{Ne}] \text{ core})$
- b) $\text{Ca} ([\text{Ar}]4s^2) \xrightarrow{-2e^-} \text{Ca}^{+2} ([\text{Ar}] \text{ core})$
- c) $\text{Co} ([\text{Ar}]4s^23d^7) \xrightarrow{-3e^-} \text{Co}^{+3} ([\text{Ar}]3d^6) \quad (\text{The } 4s \text{ electrons are lost first.})$
- d) $\text{Ag} ([\text{Kr}]5s^14d^{10}) \xrightarrow{-1e^-} \text{Ag}^+ ([\text{Kr}]4d^{10}\underline{5s^0}) \quad (\text{The } 5s \text{ electron is lost first.})$
- e) $\text{I} ([\text{Kr}]5s^24d^{10}5p^5) \xrightarrow{+1e^-} \text{I}^- ([\text{Kr}]5s^24d^{10}5p^6 \text{ or } [\text{Xe}] \text{ core})$
- f) $\text{Bi} ([\text{Xe}] 6s^24f^{14}5d^{10}6p^3) \xrightarrow{-3e^-} \text{Bi}^{+3} ([\text{Xe}] 6s^24f^{14}5d^{10}\underline{6p^0})$
- g) $\text{Sc} ([\text{Ar}]4s^23d^1) \xrightarrow{-3e^-} \text{Sc}^{+3} ([\text{Ar}] \text{ core})$
- h) $\text{P} ([\text{Ne}]3s^23p^3) \xrightarrow{+3e^-} \text{P}^{3-} ([\text{Ne}]3s^23p^6 \text{ or } [\text{Ar}] \text{ core})$

47. NH_4^+ is a cation which is not a metal ion. Likewise, $(\text{NH}_4)_2\text{SO}_4$ is a salt (ionic compound) which doesn't contain a metal and a nonmetal but rather contains only nonmetal elements.

48. Recall that the force of attraction between ions is governed by Coulomb's Law.

$$F \propto \frac{q_1 q_2}{r^2}$$

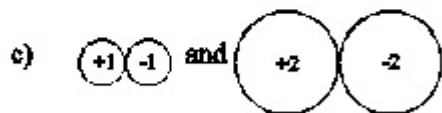
Note that when q_1 and q_2 have opposite charge the force (F) carries a negative sign (attraction) and when q_1 and q_2 have like charges the force (F) carries a positive sign (repulsion).



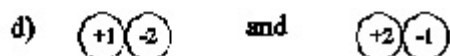
The cations and anions in each set carry the same charges so the force depends only on the sizes of the ions. In the second pair the anion is twice the size of the anion in the first pair. Therefore, the force of attraction in the first pair is twice as high as in the second pair.



The cations and anions in each set are the same size so the force depends only on the charges on the ions. In the second pair of ions the charges are twice as high as in the first pair. Thus, the force of attraction in the second pair is four times as high as in the first pair.



The charges on the ions in the second pair are twice those of the first pair. The sizes of the ions in the second pair are also twice as large as in the first pair. The effects of increased size and charge cancel making the two forces equal.



The sizes of all of the ions are the same. Therefore, the force depends only on the charges on the ions. However, $q_1 \times q_2$ is equal to -2 for each pair. Thus the forces of attraction are the same for the two pairs.

49. The following predictions are based upon interpretation of the changes in the electrostatic forces between two ions (one cation, A, and one anion, B) based on Coulomb's Law. The change in energy between each combination of ions would probably not match **QUANTITATIVELY** the changes in the ionic crystal lattice energy. Never-the-less, the electrostatic (or coulombic) changes predicted below should reflect general changes in the ionic crystal lattice energy and therefore in the general stability of the ionic lattice.

Coulomb's Law states that,
$$F \propto \frac{q_1 q_2}{r^2}$$

- a) Doubling the charge on A doubles the force
 - b) Doubling the charge on B doubles the force
 - c) Doubling the charges on both A and B quadruples the force
 - d) doubling the radii of both A and B doubles the interatomic distance reduces the force to one-fourth
50. The force of attraction in salts (ionic compounds) is governed by Coulomb's Law.
$$F \propto \frac{q_1 q_2}{r^2}$$
- a) NaF; Both NaF and NaCl contain Na^+ ions and -1 ions. Therefore any difference in the forces of attraction in the two salts must depend only on the difference in sizes of the anions involved. Since the F^- ion is smaller than the Cl^- ion then by Coulomb's Law the force of attraction between Na^+ and F^- ions must be greater than the force of attraction between Na^+ and Cl^- ions.
 - b) MgO; The size of $\text{Na}^+ \approx \text{Mg}^{+2}$ and the size of $\text{F}^- \approx \text{O}^{-2}$. Therefore, any difference in the forces of attraction in the two salts must be due to difference in the charges on the respective ions. In MgO the charges on both the cation and anion are twice as large as the charge on in cation and anion in NaF. Therefore, the force of attraction in MgO is approximately four times the force of attraction in NaF.
 - c) KCl; Both KCl and RbBr contain $+1$ and -1 ions so again any differences in the strengths of the attractions must be due to difference in size among the ions involved. Both the cation and anion are larger in RbBr than in KCl so the force of attraction in KCl must be larger than in RbBr.
51. For isoelectronic species the size should decrease as Z^* increases (or as atomic number, Z , increases).
Therefore: $\text{Y}^{+3} < \text{Sr}^{+2} < \text{Rb}^+ < \text{Kr} < \text{Br}^- < \text{Se}^{-2}$
52. Recall the trends in ionic radii:
- i) Increases down a family or ions with like charges.
 - ii) Decreases with increasing nuclear charge for isoelectronic ions.
 - iii) Cations are always smaller than their parent atom. If more than one cation can be formed from the same parent atom, ionic radius decreases with increasing charge.
 - iv) Anions are always larger than their parent atom.
- a) $\text{Br} < \text{Br}^-$
 - b) $\text{Ag} > \text{Ag}^+$
 - c) $\text{O}^{-2} > \text{F}^-$
 - d) $\text{Au}^+ > \text{Au}^{+3}$
 - e) $\text{In}^{+3} < \text{Tl}^{+3}$
 - f) $\text{In}^+ > \text{Sn}^{+2}$

53. a) $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$ b) $\text{Al}^{+3} < \text{Mg}^{+2} < \text{Na}^+$ c) $\text{Be}^{+2} < \text{Al}^{+3} < \text{Na}^+ < \text{Cs}^+$
 d) $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$ e) $\text{F}^- < \text{O}^{-2} < \text{N}^{-3}$ f) $\text{O}^{-2} < \text{S}^{-2} < \text{P}^{-3} < \text{I}^-$

QUESTIONS ABOUT BONDING

54. As the "Noble Gas Configuration" is of considerable stability, simple monatomic ions tend to form with these same electron configurations. The valence electron configurations of the "A-Family" metals (e.g., I-A and II-A metals), are such that these elements can achieve the Noble Gas Configuration by losing a relatively small number of electrons during chemical reactions. Consequently, these elements tend to form cations during chemical reactions. For example, the I-A elements have a ns^1 valence electron configuration and therefore, can achieve the Noble Gas Configuration by losing this one valence electron. It is for this reason that the I-A elements form ions with a +1 charge.
55. The valence electron configurations of the "A-Family" nonmetals (e.g., VI-A and VII-A elements), are such that these elements can achieve the Noble Gas Configuration by gaining a relatively small number of electrons during chemical reactions. Consequently, these elements tend to form anions during chemical reactions. For example, the VII-A elements have a ns^2np^5 valence electron configuration and therefore, can achieve the Noble Gas Configuration by gaining only one valence electron. It is for this reason that the VII-A elements form ions with a -1 charge.
56. $\Delta(\text{EN})$ is a measure of individual bond polarity, i.e., a greater $\Delta(\text{EN})$ produces a greater bond polarity.

a)

BOND	H—F	H—C	H—H
EN of Atoms	2.1 , 4.0	2.1 , 2.5	2.1 , 2.1
$\Delta(\text{EN})$	1.9	0.4	0.0
Polarity of Bonds	H—H < H—C < H—F		

b)

BOND	P—S	Si—Cl	Al—Cl
EN of Atoms	2.1 , 2.5	1.8 , 3.0	1.5 , 3.0
$\Delta(\text{EN})$	0.4	1.2	1.5
Polarity of Bonds	P—S < Si—Cl < Al—Cl		

c)

BOND	Cl—I	P—P	C—N	O—H
EN of Atoms	3.0 , 2.5	2.1 , 2.1	2.5 , 3.0	3.5 , 2.1
$\Delta(\text{EN})$	0.5	0.0	0.5	1.4
Polarity of Bonds	P—P < Cl—I $\approx$ C—N < O—H			

57. Recall the criteria for the prediction of bonding types:

$\Delta(\text{EN}) \geq 1.8$ means the compound is predicted to be ionic.

$\Delta(\text{EN}) \leq 1.8$ means the compound is predicted to be polar covalent.

$\Delta(\text{EN}) \approx 0.0$ means the compound is predicted to be nonpolar covalent.

a)

Compound Formula	I_2
EN of Atoms	2.5 , 2.5
$\Delta(\text{EN})$	0.0
Bond Type	Nonpolar Covalent

b)

Compound Formula	LiCl
EN of Atoms	1.0 , 3.0
$\Delta(\text{EN})$	2.0
Bond Type	Ionic

c)

Compound Formula	MgTe
EN of Atoms	1.2 , 2.1
$\Delta(\text{EN})$	0.9
Bond Type	Polar Covalent

d)

Compound Formula	Cl_2O
EN of Atoms	3.0 , 3.5
$\Delta(\text{EN})$	0.5
Bond Type	Polar Covalent

e)

Compound Formula	H_2S
EN of Atoms	2.1 , 2.5
$\Delta(\text{EN})$	0.4
Bond Type	Polar Covalent

f)

Compound Formula	Cs_2O
EN of Atoms	0.7 , 3.5
$\Delta(\text{EN})$	2.8
Bond Type	Ionic

g)

Compound Formula	BCl_3
EN of Atoms	2.0 , 3.0
$\Delta(\text{EN})$	1.0
Bond Type	Polar Covalent

h)

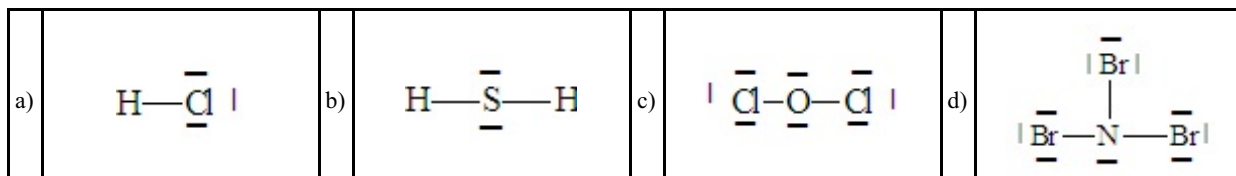
Compound Formula	SO_2
EN of Atoms	2.5 , 3.5
$\Delta(\text{EN})$	1.0
Bond Type	Polar Covalent

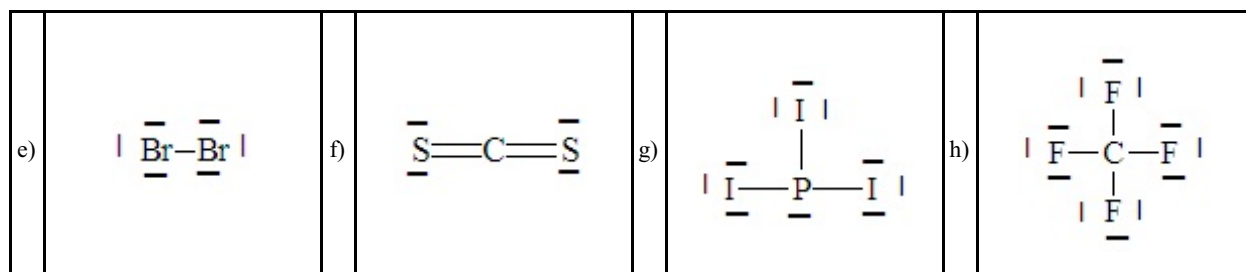
i)

Compound Formula	Al_2O_3
EN of Atoms	1.5 , 3.5
$\Delta(\text{EN})$	2.0
Bond Type	Ionic

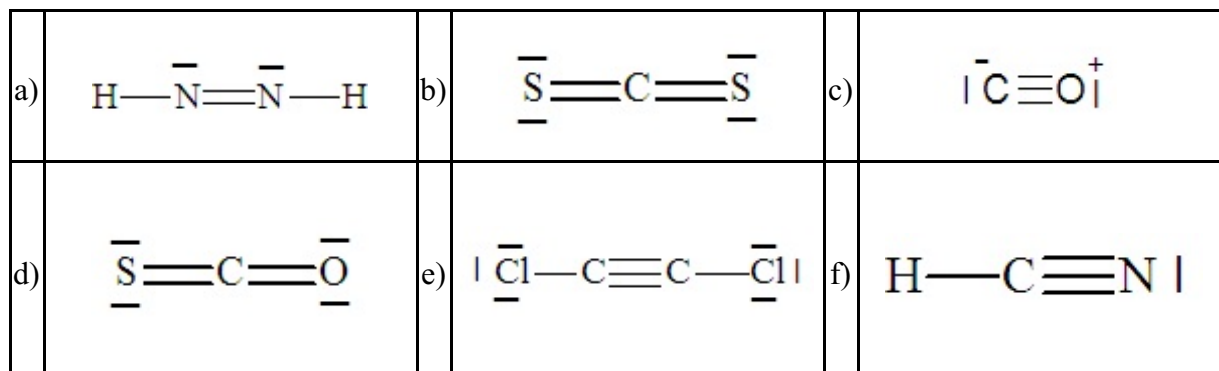
QUESTIONS ABOUT LEWIS STRUCTURES

58.

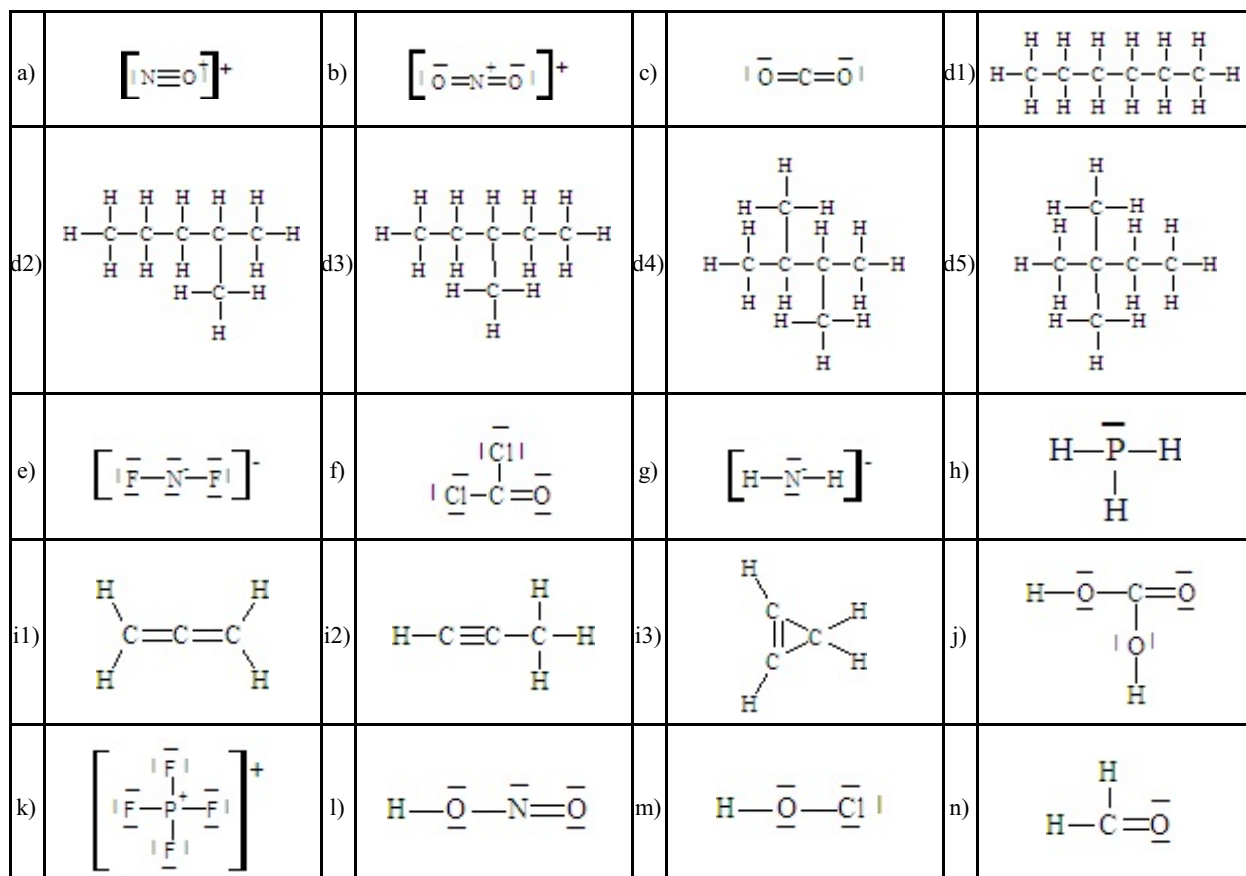




59.



60. Note: Lewis structures are not intended to depict geometry - only electron distributions.
 A * by the structure indicate that resonance is present.



o)		p)		q1)		q2)	
q3)		r)		s1)		s2)	
t1)		t2)					

61.

a1)		a2)		a3)		b1)	
b2)		c1)		c2)		c3)	
d1)		d2)		d3)		d4)	
e1)		e2)		e3)		e4)	
e5)		f)		g)		h)	

i)		j1)		j2)		j3)	
j4)		j5)		j6)		j7)	
j8)		j9)		j10)		j11)	
j12)		j13)		j14)		j15)	
j16)		k1)		k2)		l)	
m)							

62. i) a) linear on both N and O: DNA. b) linear on N: DNA.
 c) linear on C: Polar. d) tetrahedral on all C's in every isomer: Polar.
 e) angular on N: DNA. f) trigonal planar on C: Polar.
 g) angular on N: DNA. h) trigonal pyramidal on P: Polar.
 i1) trigonal planar/linear/trigonal planar: Nonpolar. i2) linear/linear/tetrahedral: Nonpolar
 i3) trigonal planar on 2 C's and tetrahedral on 1 C: Polar.
 j) trigonal planar on C and angular on O: Polar. k) tetrahedral on P: DNA.
 l) angular on both N and O: Polar. m) angular: Polar.
 n) trigonal planar: Polar. o) trigonal planar on both C: Nonpolar.
 p) trigonal planar on both C's: Polar. q1) trigonal planar on both C's: Polar
 q2) trigonal planar on both C's: Nonpolar q3) trigonal planar on both C's: Polar
 r) tetrahedral on both C's: Polar. s1 & s2) tetrahedral on both C's: Polar.
 t1 & t2) tetrahedral on both C's and angular on O: Polar.
- ii) a) angular; sp^2 on S. b) angular; sp^2 on N.
 c) trigonal planar; sp^2 on C. d) tetrahedral; sp^3 on S and angular; sp^3 on O.
 e) tetrahedral; sp^3 on P. f) trigonal bipyramidal; sp^3d on P.
 g) tee-shaped; sp^3d on I. h) linear; sp^3d on I.
 i) octahedral; sp^3d^2 on P. j) tetrahedral sp^3 on Xe.
 k) square pyramidal; sp^3d^2 on Xe. l) square planar; sp^3d^2 on I.
 m) square planar; sp^3d^2 on Xe.

iii) The structures that are polar are: **a, d, f, g, and k.**

The structures that are nonpolar are: **j, and m.**

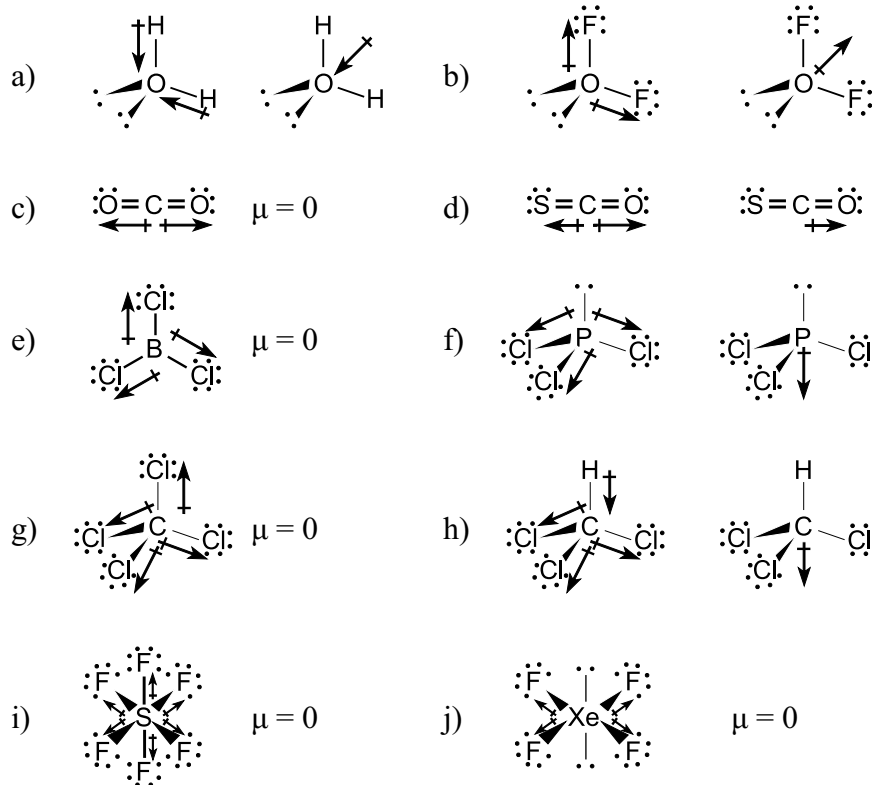
The structures in which the terms, polar and nonpolar

DO NOT APPLY because they are ions are: **b, c, e, h, i, and l.**

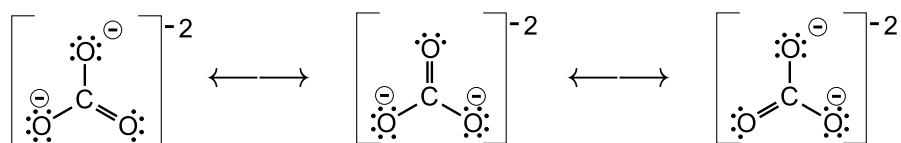
63. Since these molecules are diatomic, the bond polarity is equal to the overall molecular polarity.

BOND	I—Cl	Br—Cl	Cl—Cl	Br—F	Cl—F
EN of Atoms	2.5 , 3.0	2.8 , 3.0	3.0 , 3.0	2.8 , 4.0	3.0 , 4.0
$\Delta(EN)$	0.5	0.2	0.0	1.2	1.0
Polarity of					
Bonds or	Cl—Cl	Br—Cl	I—Cl	Cl—F	Br—F
Molecules	<	<	<	<	

64. Individual bond moments are sketched as vectors on the indicated molecular shape (FIRST STRUCTURE). The relative magnitude of the bond moments are deduced from $\Delta(\text{EN})$. If a net molecular dipole results, its direction is indicated on the SECOND STRUCTURE of the molecule.

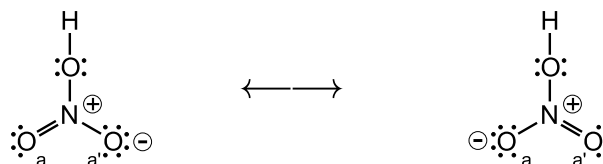


65. a) In the Lewis structure for CO_2 , $\text{:}\ddot{\text{O}}=\text{C}=\ddot{\text{O}}\text{:}$, there is a double bond between carbon and each oxygen but, in CO_3^{-2} , there are three resonance structures;



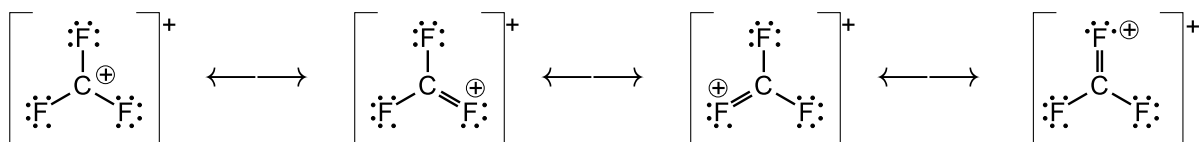
Thus, on average there are 1 and 1/3 bonds between each oxygen and the carbon in CO_3^{-2} . As bond length (or distance) decreases as the bond order (the number of bonds) increases, the bonds in CO_2 are shorter than the bonds in CO_3^{-2} .

b) There are two resonance structures for nitric acid, HNO_3 .

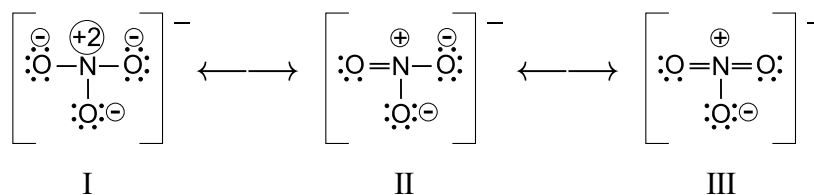


The effect of the resonance is to average out the bonding between the nitrogen and the two terminal oxygens (a and a' in the pictures) at a bond and one half each. Therefore these two bonds (a and a') are alike and are shorter than the remaining nitrogen to oxygen bond (which is only a single bond).

- c) Since the observed bond length (127 pm) in CF_3^+ is shorter than the normal C–F single bond (138 pm), it follows that the bonding in CF_3^+ must involve some multiple bonding. The **BEST** single Lewis structure for CF_3^+ is the one with only three single bonds, but, this cannot be the only structure which contributes to the true structure of this ion as this would predict a bond length of approximately 138 pm. The three structures which contain a double bond between one F and the C also contribute to the true structure. This has the effect of increasing the bonding between the three fluorines and the carbon thereby decreasing the bond lengths to the observed value.

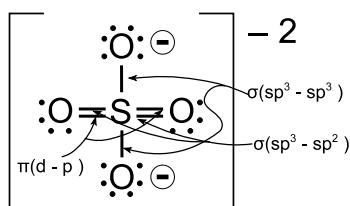


66. Formal charges for the atoms in each structure must be determined before a comparison can be made.

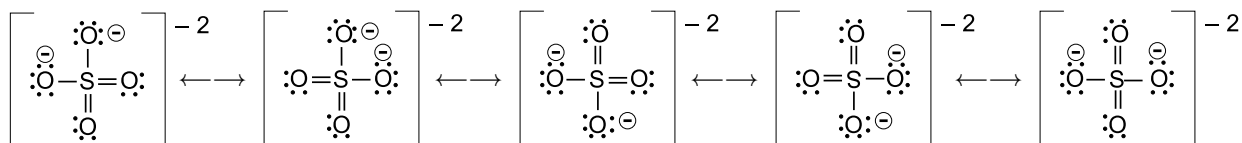


First of all, III is an impossible structure as it contains a N atom with five bonds to it. Nitrogen is a second row element and therefore has only four valence orbitals with a maximum bonding capacity of four. Structure II has more bonds and fewer formal charges than I, therefore, II is a better Lewis structure than I. Therefore, II is the best structure.

67. a)



- b) Yes, other resonance structures are possible. In fact there are 5 other resonance structures which are of equivalent energy to the above structure.



In addition there are four other resonance structures which have only one double bond in them and one structure with all single bonds. These structures are not as stable as the above six structures but, do contribute something to the overall structure for SO_4^{2-} .