

CHEM 245**ANSWERS TO PS 1:**

1. a) 2.426×10^{-11} m
 b) 6×10^{-34} m

2. $E = 5.002 \times 10^{-19}$ J ; $\lambda = 397.0$ nm

3. Bohr predicted that only certain orbits whose radii are given by a certain equation are allowed to the electron.

$$r = (n^2 / Z) a_0$$

4. See text books.

5. Locations where the radial wavefunction passes through zero are Radial Nodes.
 The planes on which angular wave function passes through zero are called Angular Nodes or Nodal Planes.

6. Radial nodes: For 3p = 1 For 4d = 1
 Angular nodes: For 3p = 1 plane For 4d = 2 planes

7. a)

	n	l	m_l
5d	5	2	-2, -1, 0, 1, 2
4f	4	3	-3, -2, -1, 0, 1, 2, 3
7g	7	4	-4, -3, -2, -1, 0, 1, 2, 3, 4

b)

	n	l	m_l	m_s
3d	3	2	-2	$\pm\frac{1}{2}$
			-1	$\pm\frac{1}{2}$
			0	$\pm\frac{1}{2}$
			1	$\pm\frac{1}{2}$
			2	$\pm\frac{1}{2}$

8. Ce: $Z_{eff} = 12$
 O^{2-} : $Z_{eff} = 3.85$
 Na^+ : $Z_{eff} = 5.85$

9. a) Ca: $Z_{eff} = 2.85$; Ga: $Z_{eff} = 5$
 b) Mg: $Z_{eff} = 2.85$; Ca: $Z_{eff} = 2.85$
 c) Si: $Z_{eff} = 4.15$; P: $Z_{eff} = 4.80$
 d) Mn^{2+} : $[Ar] 3d^5$

10. a) S_{16} : $Z_{eff} = 5.45$ EA = 2.077 eV $r = 102$ pm
 Cl_{17} : $Z_{eff} = 6.1$ EA = 3.617 eV $r = 99$ pm

Chlorine has larger Z_{eff} and smaller radius compared to sulphur. They both increase the attractive power of an electron. **Therefore $\rightarrow EA_S < EA_{Cl}$**

b) I_{53} : $Z_{eff} = 7.6$ EA = 3.059 eV
 Br_{35} : $Z_{eff} = 7.6$ EA = 3.365 eV

Large Z_{eff} and small radius \rightarrow large EA!

$r_{Br} < r_I$ **Therefore $\rightarrow EA_{Br} > EA_I$**

c) B_5 : $1s^2 2s^2 2p^1$ IE = 8.298 eV
 Be_4 : $1s^2 2s^2$ IE = 9.322 eV

Boron starts a new p subshell. It is easier to remove p^1 electron compared to s^2 electron! **Therefore $\rightarrow IE_B < IE_{Be}$**

d) S_{16} : $1s^2 2s^2 2p^6 3s^2 3p^4$ IE = 10.360 eV
 P_{15} : $1s^2 2s^2 2p^6 3s^2 3p^3$ IE = 10.486 eV

Sulfur has 4 electrons in $3p$ orbital. There will be an $e^- - e^-$ repulsion at the doubly occupied p orbital. Also P has a half filled e^- configuration, which results in an extra stability. **Therefore $\rightarrow IE_S < IE_P$**

e) Cl_{17} : $1s^2 2s^2 2p^6 3s^2 3p^5$ $Z_{eff} = 6.1$ $r = 99$ pm IE = 12.977 eV
 F_9 : $1s^2 2s^2 2p^5$ $Z_{eff} = 5.2$ $r = 71$ pm IE = 17.422 eV

Therefore $\rightarrow IE_F < IE_{Cl}$

11. a) $Se^{2-} > Br^- > Rb^+ > Sr^{2+}$

They are all isoelectronic. As Z_{eff} increases, the size decreases! The general rule is $M^- > M > M^+$

b) $Y^{3+} > Zr^{4+} > Nb^{5+}$ isoelectronic, As Z_{eff} increases, the size decreases!

c) $Co^{4+} < Co^{3+} < Co^{2+} < Co$ smaller the charge, bigger the size!

12 and 13. I am sure you can do these questions easily!

14. a. AsH_3 Largest central atom, minimized bond-pair – bond-pair repulsion.
b. O_3^- Nearer to the tetrahedral angle, almost 4 e- pairs around the central O atom.
c. F-S-F since F is more electronegative and pulls electrons towards itself.
d. ONO Electrons will be pulled away from N, reducing the repulsion between bond-pairs.
e. BrO_3^- Br is the larger atom.

15. Done all these in class and lab.

16. Bond angles increase as the attached groups become bulkier.

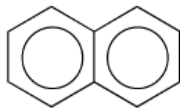
17. a) N b) O c) C d) O e) Br f) Cl

Order of polarity: $\text{O-Cl} > \text{N-O} > \text{C-N} > \text{P-Br} > \text{S-Cl} > \text{C-I}$

18. $\mu = Q \cdot r$ r = separation, Q = charge on each atom, μ = dipole moment

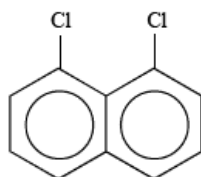
ANSWERS TO PS 2:

a) Naphthalene



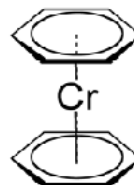
D_{2h}
Not chiral
Not polar

b) 1,8-Dichloronaphthalene



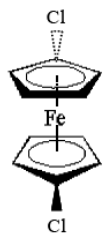
C_{2v}
Not chiral
Polar

c) Dibenzenechromium



D_{6h}
Not chiral
Not polar

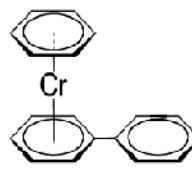
d) 1, 1'-Dichloroferrocene



Not chiral
Not polar

C_{2h}

e)



Not chiral
Polar

C_s

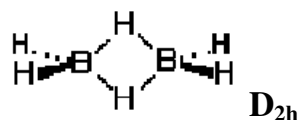
f) Cyclohexane (chair)



D_{3d}

Not chiral, Not polar

g) Diborane



D_{2h}

Not chiral, Not polar

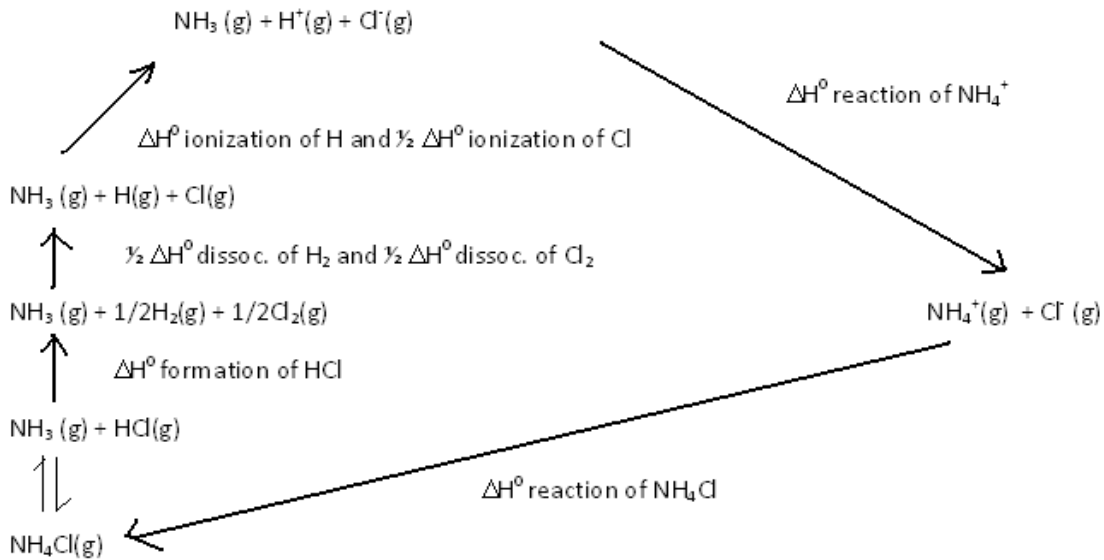
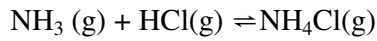
h) Number 96 (font: Times New Roman)

C_{2h} **Not chiral, Not polar**

ANSWERS TO PS 3:

1. a) close-packed b) not close-packed c) closest-packed
2. W atoms are much larger than W atoms. We can visualize WC as a close-packed (FCC) lattice of W atoms with C atoms in the octahedral holes. C.N. for both atoms = 6.
3. a) C. N. for Cs^+ = 8 b) C. N. for Cl^- = 8 c) 6
4. $a = 2 \times r_{Cl^-} + 2 \times r_{Na^+} = 552 \text{ pm}$
5. $c = 4r = (3)^{1/2}a$ $r = 124 \text{ pm}$
6. Density of Fe = 7.85 g Fe/cm^3
7. It is an example to Spinel. Possibilities are $A^{2+} B^{3+}_2$ $A^{4+} B^{2+}_2$ etc
8. $r_B / r_X = 0.414$ $a = r_X + 2r_B + r_X$

9. Take the following as the starting reaction:



ANSWERS TO PS 4:

1. $E_g \approx \Delta E = h\nu = hc / \lambda$

$$\lambda = hc / E_g = \frac{(6.626 \times 10^{-34} \text{ Js}) (3 \times 10^8 \text{ m/s})}{(2.4 \text{ eV}) (1.60 \times 10^{-19} \text{ J/eV})}$$

$\lambda = 5.18 \times 10^{-7} \text{ m} = 518 \text{ nm}$

2. Conductivity,

$$\frac{\sigma_{373}}{\sigma_{273}} = e^{-E_a / kT} = e^{6.38}$$

$\sigma_{373} / \sigma_{273} = 590$

3. In ZnO or TiO, addition of Zn or Ti would be addition of 2 more electrons with respect to the Zn^{2+} or Ti^{2+} ions. Since addition of Zn or Ti would supply extra electrons, this would result in n-type semiconductors.

In Cu_2S , CuI or ZnO, excess S, I or O would have fewer electrons than the corresponding ions (ie. S^{2-} , I^{2-} , O^{2-}). The result would be excess non-metals in the lattice, and hence p-type semiconductors.

4. a) arsenic-doped germanium: ***n*-type**
5 e⁻ 4 e⁻
- b) gallium doped germanium: ***p*-type**
3 e⁻ 4 e⁻
- c) silicon-doped germanium: **none**
4 e⁻ 4 e⁻