KEY!

Unit 3 Test Prep- Ch.7-9

1. Match the following trends to their definition
   A. Atomic Radii
   B. Electronegativity
   C. Ionization Energy
   D. Electron Affinity
   E. Ionic Radii

__A__ Half the distance between the nucleus of two identical atoms in a bond.

__C__ The energy required to remove an electron from an ion.

__E__ The distance from an ion's nucleus to its outermost electron.

__D__ The energy change when an atom gains an electron to become a negative ion.

__B__ The tendency of an atom to attract a pair of electrons.

2. Which of the trends listed above increase along the rows from right to left and along the columns from top to bottom?

A: ↓ ←

B, C, D, E: ↑ →

3. Which of the trends listed above increase along the rows from left to right and along the columns from bottom to top?

4. What is the correct order for increasing atomic radius for the elements Mg, Na, P, Si, and Ar?

Ar, P, Si, Mg, Na

5. What is an isoelectronic series? Provide an example in your explanation.

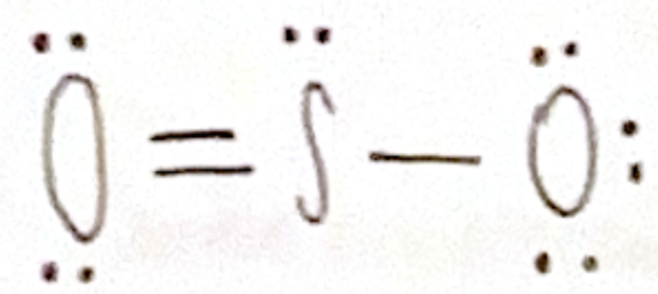
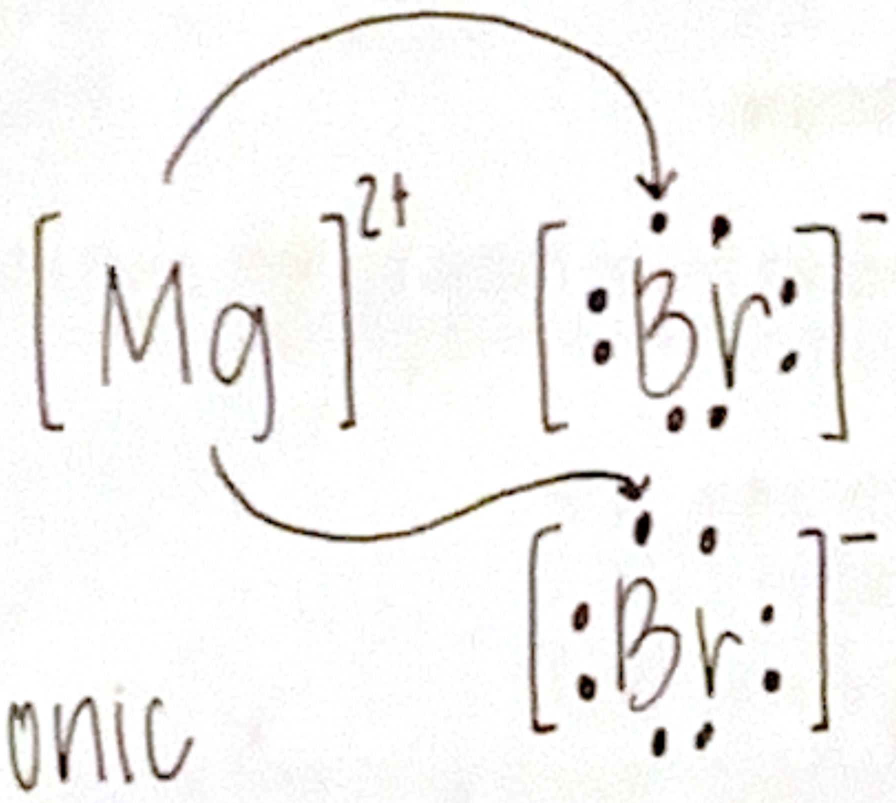
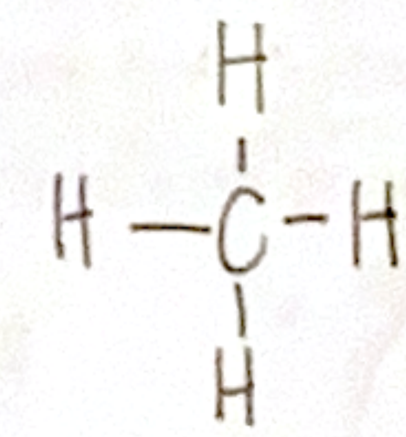
A group of ions w/ the same # of electrons and electron configuration (ex. N^{3-}, O^{2-}, F^{-}, Ne)

6. What is effective nuclear charge? Include the formula in your answer.

The net positive charge an electron experiences from the nucleus based on how far/close it is.

$$Z_{eff} = \underbrace{Z}_{\text{atomic \#}} - \underbrace{S}_{\text{shielding electrons}}$$

7. Draw the Lewis structure for the following compounds and identify if it is covalent/ionic.

SO_2 S: 1 x 6 = 6 O: 2 x 6 = 12 18 ve⁻  covalent	N_2 N: 2 x 5 = 10 ve⁻ :N≡N: covalent
$MgBr_2$  Ionic	CH_4 C: 1 x 4 = 4 ve H: 4 x 1 = 4 8 ve⁻  covalent

8. Write the formula for calculating the formal charge of an atom in a covalent Lewis structure.

$FC = ve^- - \frac{1}{2} be^- - nbe^-$

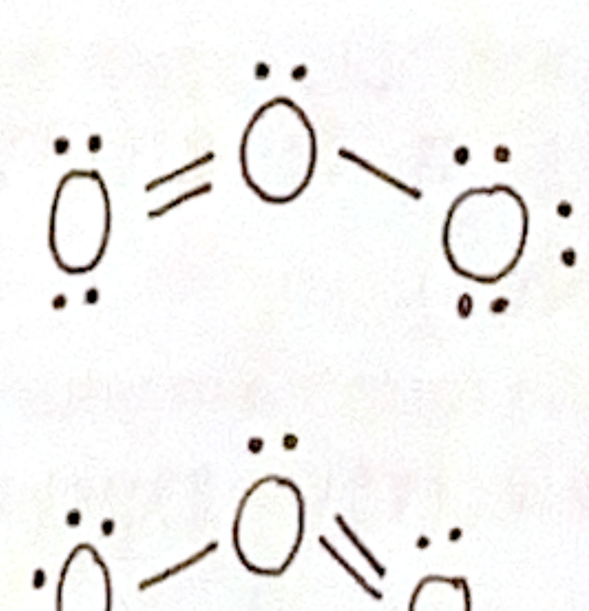
$\boxed{FC = ve^- - \text{lines} - \text{dots}}$

9. What is resonance?

Presence of delocalized electrons that can not be represented by a distinct bond. Results in multiple ways to draw a Lewis structure with varied positions of the multiple bond.

10. Determine which of the following compounds has resonance and draw all possible configurations.

a. CO_2
b. H_2O
c. O_3
d. NH_3



(creates a dipole moment) (do the dipole moments cancel out or add up)

11. Describe the difference between bond polarity and molecular polarity.

represents the sharing of electrons between 2 atoms

represents the overall distribution of charge in the molecule

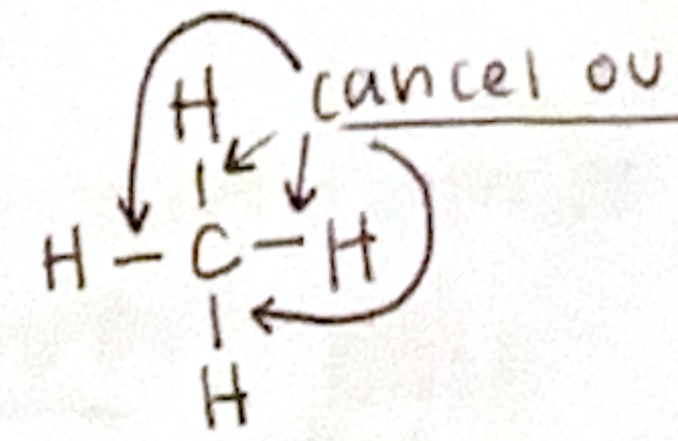
12. What is the mathematical scale used to determine polarity when calculating the difference in electronegativity between 2 atoms?

non-polar | polar | ionic

0.5 | 1.7+

13. Why can a molecule contain polar bonds but still be nonpolar overall? Give an example of a compound that demonstrates this idea.

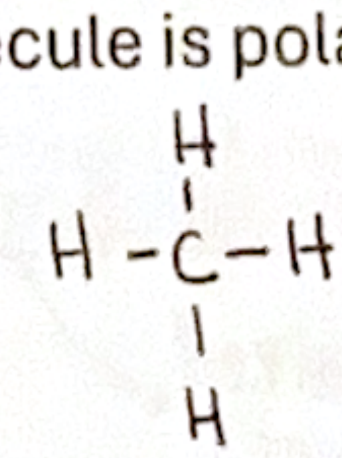
The polar bonds can create dipole moments that cancel out.

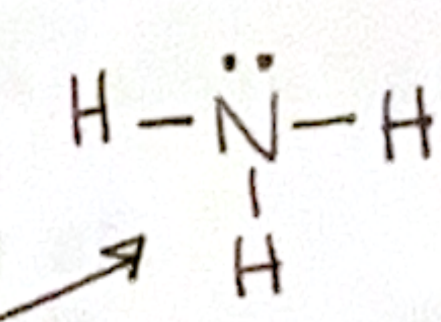


14. How does symmetry affect whether a molecule is polar or nonpolar?

symetrical = non-polar

asymetrical = polar





15. Which of the following bonds would be hardest to break due to its polarity and bond strength? (Electronegative values: H=2.1, C=2.5, N=3.0, O=3.5, F=4.0)

a. C-H $|2.1 - 2.5| = 0.4$

b. N-H $|3.0 - 2.1| = 0.9$

c. O-H $|3.5 - 2.1| = 1.4$

d. F-F $|4.0 - 4.0| = 0$

16. Explain the difference between electron geometry and molecular geometry.

electron spatial arrangement

3D shape

17. What is the molecular geometry for the following compounds based on their descriptions?

a. 3 single bonds; 1 lone pair trigonal bipyramidal

b. 1 bond; no lone pairs linear

c. 1 single bond & 1 double bond; 1 lone pair ~~linear~~ bent

d. 2 double bonds; no lone pairs linear

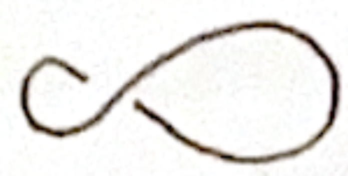
e. 4 single bonds; no lone pairs tetrahedral

18. What is a dipole moment?

A point of unevenly distributed charge from the unequal sharing of electrons between 2 atoms.

19. Draw an s, p, and sp orbital, then draw a visual example of a sigma and pi bond.

s: 

sp: 

p: 

sigma bond: 
head-to-head overlap

pi bond: 
side-to-side overlap

20. What is the formula for calculating bond order and what does it conclude?

$$\text{Bond order} = \frac{\#\,b\bar{e} - \#\,\text{anti-}b\bar{e}}{2}$$

How stable a compound is (0 = unstable, 1 = weak, 2 = double bond, 3 = triple bond)

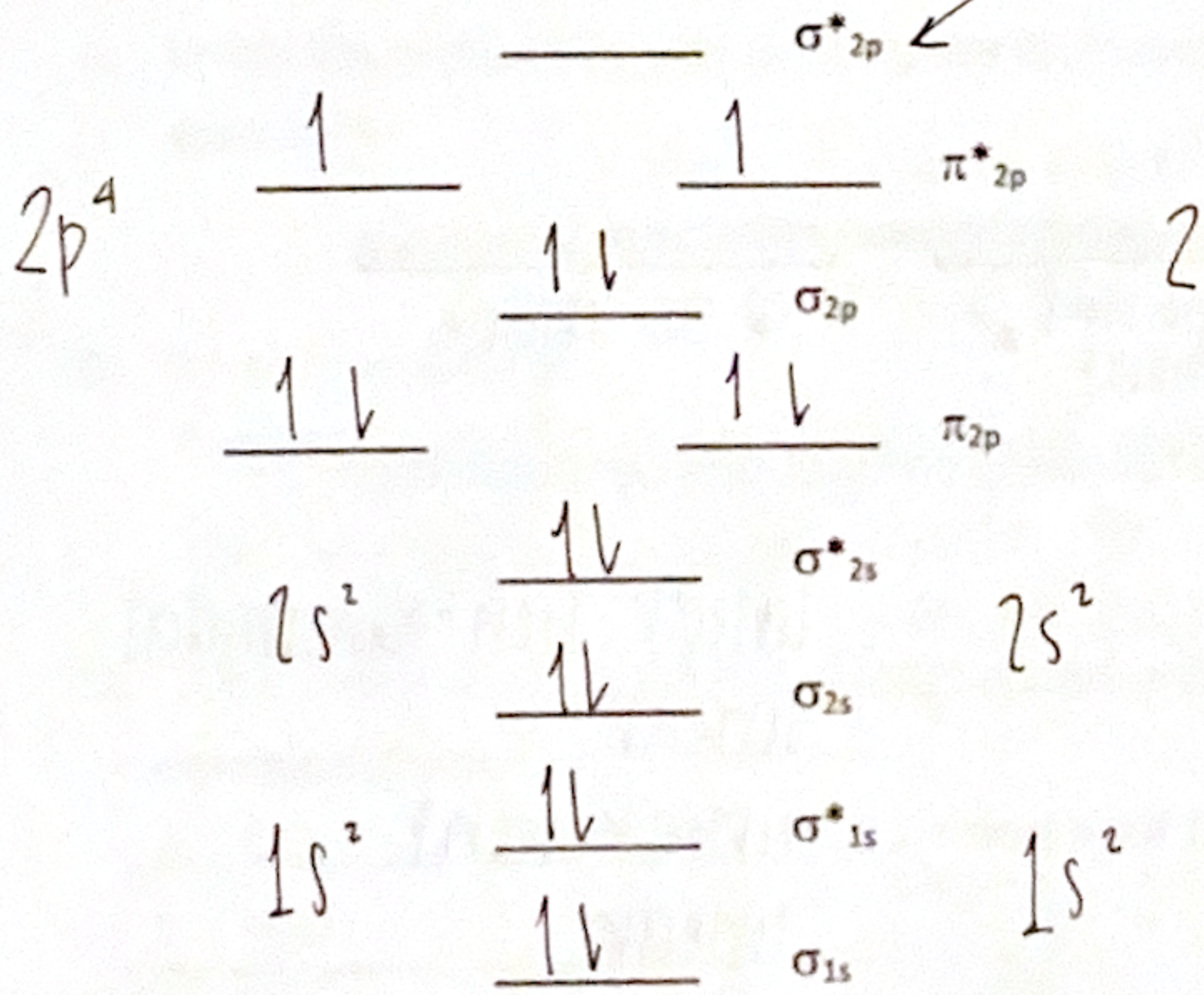
21. Explain the difference between diamagnetism and paramagnetism.

no unpaired electrons; repelled by magnets

at least 1 unpaired electron; attracted by magnets

22. Complete the molecular orbital diagram for O_2 and calculate the bond order. What is its magnetism?

unpaired electron = paramagnetic



$$\text{Bond order} = \frac{8-2}{2} = \boxed{2}$$