17. What is meant by an excited state? Give the electron configuration for a possible excited state for an atom of Li.

Excited state: an electron jumps up to a higher energy level than its ground state

ex: Li ground state: $1s^2 2s^1$ Li excited state: $1s^2 2p^1$

18. How many **orbitals** occur in a set of p-orbitals? 3

19. How many **electrons** can fit into a set of d-orbitals? 10

20. Predict the electron configuration for each of the given species (core notation is okay!):
   a) Mo: [Kr] $5s^2 4d^4$ ⇒ [Kr] $5s^1 4d^5$
   b) S^{2-} [Ne] $3s^2 3p^6$
   c) In^{3+} [Kr] $4d^{10}$

21. Write the electron configurations for each of the given species (Do not use core notation!)
   a) The phosphide ion: $1s^2 2s^2 2p^6 3s^2 3p^6$
   b) The indium (III) ion: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$
   c) The nickel (II) ion: $1s^2 2s^2 2p^6 3s^2 3p^6$ $\cancel{4s^2}$ $3d^8$
   d) The (very rare) Br^{3+} ion: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

22. What element has the configuration [Rn] $7s^2 5f^3 6d^1$? U

Section 2: The Periodic Table

23. Elements in Group 17 have similar chemical properties. Explain why they have such similarities, using actual electron configurations to support your answer.

All elements in group 17 want one more valence e^- to complete their valence shell.

ex: F: $1s^2 2s^2 2p^5$ & Br: [Ar] $4s^2 3d^{10} 4p^5$

24. Identify the following elements, families or parts of the periodic table. Use symbols for specific elements; where applicable, use the correct name for families and parts of the periodic table.
   a) 2^{nd} period halogen F
   b) Highest ionization energy in group 15 N
   c) Element with 6 outer s and p electrons in 4^{th} period Se
   d) Smallest atom in 5^{th} period Xe
   e) Weakest attraction for electrons in 3^{rd} period Na
   f) Family whose atoms have two weakly held electrons Alkaline Earth Metals
   g) Part of the periodic table where d-orbitals are filled Group 12
   h) Element X is in the 4^{th} period. It combines with F to form XF_3. X is: Ga

25. Why do atomic radii increase going down a group in the periodic table?

Extra shells are added & shielding means that the valence e⁻s feel the same nuclear charge

26. Why do atomic radii decrease from left to right across a period in the periodic table?

Nuclear charge increases & no new shells are added

27. Using C and Si to illustrate your answer, show your understanding of the terms *nuclear charge* and *shielding*:

   a) State the nuclear charge of C and Si C: +6 Si: +14
   b) State the # of shielding electrons for C and for Si C: 2 Si: 10

28. What observation does the ***shielding effect*** explain?

Atomic radius increases down a group

29. Which of the following pairs has the greatest attraction for its outer electrons? Briefly explain your choice..

   a) K or Cs K because smaller radius
   b) Br or Kr Kr because smaller radius
   c) Na or Ne Ne because full valence shell
   d) Mg^{+2} or Mg Mg^{+2} because smaller radius

30. A potassium atom readily loses an electron to form a positive ion. How does the size of the ion compare with the neutral atom? Explain why the size changes as it does.

Smaller → one e^- removed, which means less e^- repulsion, so smaller size

31. An oxygen atom will gain two electrons and form an O^{-2} ion.

   a) Give the electron configuration for an oxygen atom: $1s^2 2s^2 2p^4$
   b) Why does O gain two electrons?

   if it gets 2 electrons, it will have a full valence shell

   c) How does the size of the O^{-2} ion compare with the neutral atom? Why?

   bigger → 2 e^-s added, which means more e^- repulsion, so bigger size

32. Consider the neutral atoms of I, Xe, Cs, and Ba. Which element has:

   d) The largest ionization energy? Xe
   e) The largest atomic radius? Cs

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33. The ions Ca^{2+} and S^{2-} have the same number of electrons: Which one is larger? Explain.

S^{-2} ⇒ less nuclear charge, so e^-s are less attracted to the nucleus

34. a) State the two directions within the periodic table in which the ionization energies are increasing:

Circle the correct number: 1) → and ↓ 2) → and ↑ 3) ← and ↓ 4) ← and ↑

b) What explains this trend?

The smaller the atomic radius, the larger the IE

35. Define ***electronegativity***: The ability of an atom to attract neighbouring electrons

36. Which of P and Ne has the greater electronegativity? Explain why.

P ⇒ Ne has a full valence shell, and will not attract neighbouring electrons

Section 3: Chemical Bonding

37. What is an open shell?

A shell that contains less than the maximum # of e^-s

38. a) What are ***valence electrons***? # electrons in open shell

b) How many valence electrons does phosphorus have? 5

39. a) What is the ***valence*** of an atom? # unpaired electrons in valence shell

b) What is the valence of phosphorus? 3

40. What is the charge on a…

a) strontium ion? +2 b) strontium nucleus? +38 c) strontium atom? 0

41. Write the formulae of the ions which make up sodium selenide. Na^+ and Se^{-2}

42. Define ionic, covalent and polar covalent bonds. Include an explanation of what determines the nature of a given bond.

Ionic: transfer of e^-s $\Delta\chi \geq 1.7$

Covalent: equal sharing of e^-s $\Delta\chi \leq 0.2$

Polar Covalent: unequal sharing of e^-s $0.2 < \Delta\chi < 1.7$

43. What is the difference between a single, double and triple bond? Which is stronger? Explain.

single bond: 2e⁻s shared weakest
double bond: 4e⁻s shared ↓
triple bond: 6e⁻s shared strongest

44. What is the predicted formula and classification (*ionic* or *covalent*) for each of the following pairings?

Mg^{+2} P^{-3}

		Formula	***Classification***
a) Magnesium and phosphorus		Mg_3P_2	ionic
b) Carbon and iodine	⇒ valence: C=4 I=1	CI_4	covalent
c) Ca^{2+} and PO_4^{3-}		$Ca_3(PO_4)_2$	ionic
d) Arsenic and oxygen	As=3 O=2	As_2O_3	covalent

45. Of $MgCl_2$ and SrI_2, which would you predict should have the higher melting point? Why?

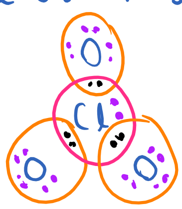
$MgCl_2$ ⇒ smaller ionic radius, so smaller distance between the ions ⇒ larger force holding them together

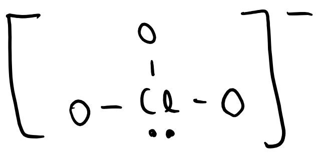
46. Of NaCl and MgS, which would you predict should have the higher melting point? Why?

MgS ⇒ similar sizes, but MgS has larger charge ⇒ larger force

47. Draw the Lewis structure for the ~~nitrate~~ ClO_3^- chlorate anion. Describe the structural pair geometry and the molecular shape.

valence e⁻s: 7+3(6)+1 = 26e⁻s



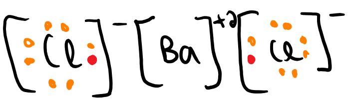


SPG: Tetrahedral

MS: Trigonal Pyramidal

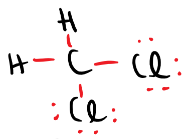
48. Draw the Lewis structures for each of the following compounds (***show your e⁻ counts***):

a) $BaCl_2$ ionic



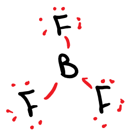
b) CH_2Cl_2

e⁻ = 4+2(1)+2(7) = 20e⁻s



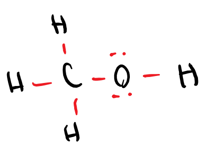
e) BF_3

e⁻ = 3+3(7) = 24 e⁻s



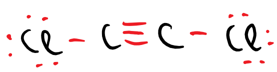
c) CH_4O

e⁻s = 4+4(1)+6 = 14e⁻s



d) C_2Cl_2

e⁻s = 2(4)+2(7) = 22e⁻s



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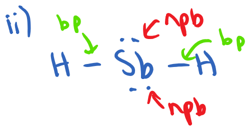
49. What are the two requirements necessary for a molecule to be polar?

1. Polar covalent bonds
2. Asymmetrical shape

50. Given SbH_2^-, **i)** draw the Lewis structure,
ii) label the bonding pairs (***bps***) and non-bonding pairs (***nbps***),
iii) state the structural pair geometry,
iv) state the molecular shape
v) identify as ***polar*** or ***non-polar***.

i) # e^-s = 5+2(1)+1 = 8

ii)



iii) SPG = tetrahedral

iv) MS = bent/angular

v) $\Delta X = 2.1 - 1.9 = 0.2$

$\Rightarrow$ non-polar

51. What type of forces are dipole-dipole and London forces? Explain the difference between them.

Intermolecular Forces ① dipole-dipole only occurs in polar molecules as a result of their polar bonds ② london forces occur in all molecules, due to instantaneous dipoles, but are only dominant in non-polar molecules

52. Predict the expected formula for the compound formed when each type of the given atoms combine based on valence. Classify the bonds as covalent, polar covalent, or ionic.

	a) ~~Ga~~ Ca and ~~S~~ F	b) **C** and ~~At~~ I	c) **Fr** and **Se**	d) **N** and **F**
Formula:	CaF_2	CI_4	Fr_2Se	NF_3
Bond Type:	ionic	covalent	ionic	polar covalent
	$\Delta X = 4.0 - 1.0 = 3.0$	$\Delta X = 2.5 - 2.5 = 0.0$	$\Delta X = 2.4 - 0.7 = 1.7$	$\Delta X = 4.0 - 3.0 = 1.0$