

Structures and Properties of Matter Review

1. In what ways were Dalton's and Thomson's atomic model similar? In what ways were they different?

In both models, atoms were spherical and solid (no empty space) and were neutral-particles. As well, in both models, atoms were considered the building blocks of all matter. Thomson's model contained internal structures and discrete structures that were smaller than the atom itself. Specifically, Thomson's model had negatively charged subatomic particles (electrons) embedded in a continuous mass of positively charged matter. Dalton's atom was the smallest possible particle of matter. It was indivisible and he believed that there were no particles smaller than the atom.

2. In what ways were Thomson's and Rutherford's models of the atom similar? In what ways were they different?

Thomson's and Rutherford's models of the atom were spherical and contained a subatomic particle believed to be the same particle in all matter (the electron). Both models included equal amounts of positively and negatively charged matter to create an overall neutral atom. In Rutherford's model, the positively charged matter was contained in a very small region at the centre of the atom which he called the nucleus. Most of the matter of the atom was contained within this nucleus. Negatively charged electrons orbited the nucleus, in a largely empty space. Thomson's model suggested that the positive and negatively charged matter was evenly distributed and no empty space existed within the atom.

3. What is common to all atomic models, from Dalton's to Bohr's?

In all atomic models, the atom is the building block of all matter, atoms are neutrally charged, and atoms are spherical.

4. Describe how the motion of an electron in Bohr's model of the atom is different from the motion of an electron in the quantum mechanical model of the atom.

Bohr's electron moves in a circular pathway a fixed distance from the nucleus; therefore, its position is known. The electron in the quantum mechanical model moves constantly within a region of space that is described by using wave equations. Its exact position cannot be known for certain.

5. Explain how periodic trends in ionization energy and electron affinity explain why atoms of elements in Group 1 and Group 2 tend to bond with other elements by forming positive ions in ionic compounds.

Groups 1 and 2 metals have low electron affinity and low ionization energies and therefore lose electrons easily, forming positive ions. These positive ions are then attracted to negative ions to form ionic compounds.

6. A nitrogen atom has a total of three electrons in its 2p orbitals. Are any of these electrons paired? Explain how you know.

There are three equal-energy p orbitals. According to Hund's Rule, single electrons with the same spin must occupy each of these orbitals before additional electrons with opposite spins can occupy the same orbitals. When the three electrons in the p orbitals of nitrogen have occupied the p orbitals, there are no additional electrons to pair with the single electrons in the p orbitals.

7. What information do the quantum numbers provide about an orbital?

Quantum numbers provide information about the energy, size, shape, spatial orientation, and the relative direction of the axis of the electron of the orbital.

8. What two things does the principal quantum number (n) describe about an orbital and the electron in it?

The principle quantum number provides information about the energy and size of the orbital.

9. What is the Pauli exclusion principle and how is it related to the spin quantum number (m_s)?

The Pauli Exclusion Principle states that only two electrons of opposite spin could occupy an orbital. Each electron in an electron pair within the same orbital are identical but have opposite spins – the spin quantum number is the same number for any electron, but the sign is either + (spin up) or – (spin down).

10. What is Hund's rule and its significance?

Hund's Rule states that the lowest energy state for an atom has the maximum number of unpaired electrons allowed by the Pauli Exclusion Principle in a given energy sublevel. This rule can also be stated as single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. This rule determines the order of the filling of orbitals and is referred to as the "Bus Seat Rule".

11. Why does it make sense that electron configurations represent atoms in their ground state and not their many excited states?

An atom's chemical properties are mainly associated with its ground state electron configuration as there are a tremendous number of possible excited states. It also shows a clear pattern as you determine the electron configuration of elements in periods and blocks of the periodic table.

12. What are the general trends for atomic radius down a group and across a period in the periodic table?

The atomic radius decreases across a period (from left to right) and increases down a group. The reason that it decreases across a period is because the amount of charge in the nucleus becomes greater, thus the attractive force on the electrons becomes greater, pulling the electrons closer to the nucleus. The radius increases down a group because each period has an additional shell and thus electrons are farther from the nucleus.

13. Without consulting a periodic table, determine to which period, group, and block the element with the electron configuration $[\text{Xe}]6s^25d^{10}$ would belong. Check your answer by using your periodic table.

It would be in Period 6, d block, Group 12. Mercury.

14. Write the complete electron configuration for the element in Period 5 and Group 15.

$[\text{Kr}]5s^24d^{10}5p^3$

15. List the 3s, 5s, 2p, 4p, 3d, and 5d orbitals in order, from the lowest energy to the highest energy.

3s, 3d, 4p, 5s, 5d

16. Which element is defined by:

- $[\text{Ar}]4s^23d^8$? *Nickel (Ni)*
- $[\text{Ne}]2s^22p^1$? *Aluminum (Al)*
- $[\text{Kr}]5s^24d^{10}5p^4$? *Tellurium (Te)*

17. Referring only to a periodic table, arrange the following sets of elements in order of increasing first ionization energy. Briefly explain your reasoning.

- Na, Si, Ar *Na, Si, Ar – first ionization energy tends to increase across a period (all in Period 3)*
- Mg, Ca, Ba *Ba, Ca, Mg – first ionization energy tends to decrease down a group (all Group 2)*
- He, Li, Be *Li, Be, He – He is highest b/c it is furthest right in Period 1; Li is lowest b/c it is to the left of Be in Period 2*

18. Consider the electron configuration: $1s^22s^22p^4$.

- Assume the electron configuration represents a neutral atom in its ground state. What element does it represent? Explain how you know. *Oxygen – there are 8 total electrons; if neutral, there would also be 8 protons*
- What information does this electron configuration notation provide? *Indicates the values of the quantum numbers n and l*
- What information does this electron configuration notation not provide? *Does not indicate information regarding the magnetic and spin quantum numbers*

19. Electrons are said to be “delocalized” in metallic bonds.

- In your own words, describe what this means, and how it is different from the locations of valence electrons in ionic and covalent bonds.

Delocalized means that the electrons that originate from one atom can move away from that atom and become a part of a larger group of shared electrons (the “sea” or “pool” of electrons surrounding cluster of metal cations). In ionic bonds, the electrons are transferred specifically from one to another atom and held there to form ions. In covalent bonds, an electron from one atom pairs up with an electron from another atom and is held in a location near those two nuclei (shared).

- b. How does the “delocalization” of electrons in metallic solids explain “malleability”, the property unique to all metals?

Metallic bonds can be viewed as non-directional. The positive ions, layered in a very organized pattern, will slide over one another when stress (such as a pounding hammer) is applied, and not break the pattern. Throughout the impact, the pool of delocalized electrons helps to keep the metal together by continuing to exert a uniform attraction on the positive ions.

20. Explain how single bonds form, based on quantum mechanical concepts.

Single bonds form between two atoms when unfilled orbitals overlap along a line between the two nuclei. The electrons are then shared between the two atoms. According to molecular orbital theory, when two atomic orbitals overlap, a new type of orbital is formed. The sharing of the electrons in this new molecular orbital constitutes the bond.

21. What experimentally observed property of methane makes it necessary to invoke the concept of hybridization to explain the structure of methane?

All four of the C-H bonds in methane are identical in length and strength. Thus, the four orbitals of the carbon that have overlapped with an orbital in the hydrogen atom have to be identical – they cannot consist of one s and three p orbitals.

22. List five possible shapes of hybrid orbitals.

Linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral

23. Describe two ways in which a non-polar molecule can temporarily become a dipole.

A non-polar molecule can move close to an ion or a dipole. The charge on the ion or the end of the dipole will distort the electron distribution of the once non-polar molecule, making it a temporary dipole.

24. Explain why symmetrical molecules are non-polar and asymmetrical molecules can be polar.

In symmetrical molecules, any polar bonds are aligned so that the polarities of the bonds cancel, leaving the molecule non-polar (so long as the bonded atoms are the same – i.e., CH₄ vs CH₃Cl). If an asymmetrical molecule has any polar bonds, the asymmetry will prevent them from cancelling each other.

25. Classify the following bonds as mostly ionic, polar covalent, or mostly covalent by looking at the location of the elements on the periodic table. Check your classifications by calculating ΔEN for each.

- Li-Cl *Ionic - $\Delta EN = 2.0$*
- S-S *Covalent - $\Delta EN = 0.0$*
- C-N *Polar Covalent - $\Delta EN = 0.5$*
- Na-O *Ionic - $\Delta EN = 2.6$*

26. Use the periodic table to help you write the condensed electron configuration for the following elements and their most likely ion (if there is one):

- a. Lithium $[He]2s^1$ $Li^{1+}: 2s^2$
- b. Argon $[Ne]3s^23p^6$ *no likely ion*
- c. Chlorine $[Ne]3s^23p^5$ $Cl^{1-}: [Ne]3s^23p^6$
- d. Phosphorus $[Ne]3s^23p^3$ $P^{3-}: [Ne]3s^23p^6$

27. Use VSEPR theory to identify the electron group arrangement (VSEPR shape), the molecular shape, and the bond angle of the following molecules whose central atoms have:

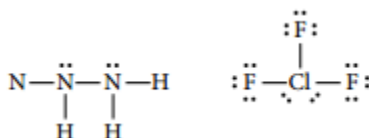
- a. 4 bonding pairs and 1 lone pair *trigonal bipyramidal* *seesaw* $90'/120'$
- b. 6 bonding pairs and 0 lone pairs *octahedral* *octahedral* $90'$
- c. 3 bonding pairs and 2 lone pairs *trigonal bipyramidal* *T-shape* $<90'$
- d. 3 bonding pairs and 0 lone pairs *trigonal planar* *trigonal planar* $120'$
- e. 2 bonding pairs and 2 lone pairs *tetrahedral* *bent* $<109.5'$
- f. 4 bonding pairs and 2 lone pairs *octahedral* *square planar* $90'$

28. Determine whether each of the following compounds will be polar or non-polar.

- a. CO_2 *nonpolar (it's linear)*
- b. H_2S *polar (it's bent)*
- c. SiO_2 *polar (it's bent)*
- d. PCl_3 *polar (trigonal pyramidal)*

29. In the 1950s, the reaction of hydrazine (N_2H_4) with chlorine trifluoride (ClF_3) was used as a rocket fuel.

- a. Draw the Lewis structures for hydrazine and chlorine trifluoride.
- b. Identify the hybrid orbitals used in each one.
- c. Identify the molecular shape and polarity of chlorine trifluoride.



hydrazine has a tetrahedral geometry, thus sp^3 ; chlorine trifluoride has a trigonal bipyramidal geometry, thus sp^3d . Chlorine trifluoride would have a T-shaped molecular shape, and would therefore be polar.

30. Use VSEPR theory and Lewis structures to predict the number on bonding pairs and lone pairs around the central atom so that you can identify the VSEPR shape, molecular shape, and bond angle for the following molecules and ions:

- a. XeF_2 *2 bonding & 3 lone pairs* *linear* $180'$
- b. BCl_4^- *4 bonding & no lone pairs* *tetrahedral* $109.5'$
- c. SF_5^+ *5 bonding & no lone pairs* *trigonal bipyramidal* $90'/120'$

31. Which compound in each of the following pairs has the higher boiling point? Explain your choice in each case.

- NH_3 or PH_3 *NH_3 – hydrogen bonding; polar molecule (P-H bonds are non-polar)*
- C_2H_6 or C_4H_{10} *C_4H_{10} – larger/more electrons, so a greater number of dispersion forces can form between molecules*
- SeCl_4 or SiCl_4 *SeCl_4 – has a seesaw shape (polar) and allows for dipole-dipole forces between molecules. SiCl_4 is tetrahedral/symmetrical, and thus, nonpolar*

32. For each of the following elements or compounds, predict which would have the higher boiling point and explain how you made your choice.

- O_2 or N_2 *O_2 – oxygen has more electrons, causing the dispersion forces to be greater*
- Ethanol or methoxymethane *ethanol – the $-\text{OH}$ group can form hydrogen bonds*
- Heptane or 2,4-dimethylpentane *heptane – larger surface area that can come in contact with other molecules; thus, experience stronger dispersion forces*