Review for Chapter 3: Atoms, Electrons and Periodic Trends

Text review questions: p. 159 – 161: 1 – 4, 6, 7, 8, 9, 13, 14, 15, 16a,b,c,d,e,h,i,j, 20, 21

1. Rutherford shot positively charged alpha particles (helium nuclei) through a piece of gold foil only a few atoms thick. He surrounded the gold foil by a zinc sulfide screen which would "light up" when hit by an alpha particle.

- the majority of alpha particles traveled straight through the gold foil. They weren't hitting anything and being stopped, so most of the gold foil must be empty space
- some of the positively charged alpha particles were slightly deflected by the foil, which meant they were being repelled by a positive charge in the gold foil that was separated from the negative charges
- some of the alpha particles were deflected right back by the foil, which meant they were hitting something very hard, dense and positively charged
- Rutherford concluded that the positive charges in the gold atoms must be concentrated in a solid, dense massive core which he called the nucleus. The negative charges were found in a random electron cloud around the nucleus, and this made up most of the volume of the atom

2.

Rutherford's Electron Cloud Model of the Atom	Bohr's Planetary Model of the Atom
• atom has a dense positively charged nucleus • electrons are found around the nucleus in a random electron cloud • electrons are found at any distance from the nucleus, with no pattern	• atom has a dense positively charged nucleus • electrons are found around the nucleus in fixed, discrete energy levels or shells • electrons are at only specific distances from the nucleus, they can move between the energy levels but can not stay in the space between levels

Bohr based his model on his study of the hydrogen bright line spectrum. When hydrogen atoms are excited by the addition of energy, their electrons absorb energy and move out to positions of higher potential energy further from the nucleus. Eventually they drop back closer to the nucleus and release their extra energy as light. Bohr passed the light through a spectroscope and found that the light had only certain colours, which meant that it had certain specific amounts of energy. He interpreted this to mean that electrons could only move between certain energy levels and were NOT in a random cloud around the nucleus. If electrons were random, they would have produced all colours of light in a continuous spectrum. Bohr revised Rutherford's electron cloud model to the planetary model, in which electrons orbit the nucleus in fixed defined paths.

3. Both Rutherford and Bohr's models of the atom have the electrons circling the nucleus, in a similar way to how our planets move around the sun, so both models have similarities to the motion of the planets. However, the planets orbit the sun in fixed, predictable defined paths. In Rutherford's model, the electrons have no fixed path so Rutherford's model is not truly a planetary model. Bohr's model does describe electrons as following fixed predictable paths, so it is more correctly a planetary model.

4a) Planck suggested that the energy of atoms is quantized; that is, it can only absorb or emit energy in certain amounts. This suggested that energy had particle-like properties. He called the smallest amount of energy that could be given off or absorbed a "quantum". So, an atom could absorb one quantum, or two quanta or three quanta of energy, but it could not absorb a half a quantum.

4b) DeBroglie suggested that if energy had particle-like properties (as proposed by Planck, above), then particles might have wave-like properties. The wave-like behaviour of small objects like electrons is significant and means that electrons probably do not move in fixed, perfectly predictable paths as Bohr suggested.

4c) Einstein supported Planck's idea that energy had particle-like properties and was quantized, so energy came in certain amounts. Einstein found that light (a form of energy) is also quantized and suggested that the packages of energy were actually photons (particles) of light.

4d) Heisenberg stated the Uncertainty Principle and showed mathematically that it is impossible to know both the position (energy) of an electron and its momentum (where it is going) at the same time. This meant that Bohr's model of the atom with the electrons moving in fixed, defined, predictable orbits could not be correct.

4e) Schrodinger pulled together the ideas of all of the above researchers. He said that if electrons have wave-like behaviour and can only exist in certain specific energy states, then the best we can do is predict where the electrons will be found most of the time. He developed mathematical equations called wave functions to allow him to predict the space around the nucleus where the electrons could be found 95% of the time, and he called these spaces orbitals.

6. Pauli's exclusion principle states that a maximum of two electrons with opposite spin can occupy the same orbital (region). This means that once an orbital contains two electrons, it is full and additional electrons will be found in other orbitals.

   Hund's rule states that electrons will only double up in an orbital once all of the orbitals in that sub-level are half full (each contain one electron). This helps to tell us the order in which to fill orbitals. For example, when filling the "p" orbitals:

   - the first electron will be an up-spin electron in the p_x orbital,
   - the second electron will be an up-spin electron in the p_y orbital,
   - the third electron will be an up-spin electron in the p_z orbital,
   - the fourth electron will be a down-spin electron in the p_x orbital and so on

Questions 7, 8, and 9, see answers in text on page 574

13. "c" is the correct orbital diagram for the third and fourth energy levels of vanadium.

    The electron configuration for vanadium is $[Ar]4s^2\ 3d^3$. The 4s orbital will be full, and the first three 3d orbitals will each contain one electron with an up-spin.

14a) carbon's electron configuration should be written $1s^2\ 2s^2\ 2p^2$ or $\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow \mid \uparrow \mid \quad}$

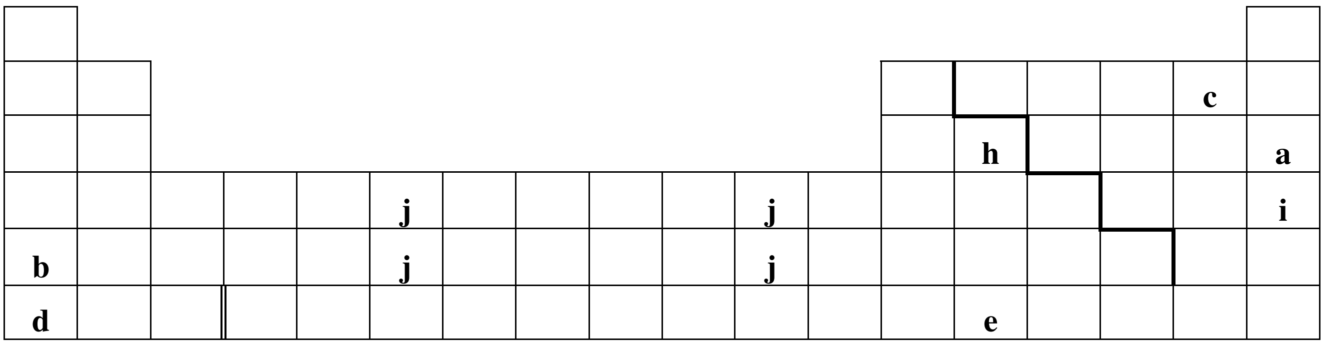
14b) iron's electron configuration is $[Ar]\ 4s^2\ 3d^6$ or $\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow \mid \uparrow \mid \uparrow \mid \uparrow \mid \uparrow}\ \boxed{\quad \mid \quad \mid \quad}$

14c) bromine's electron configuration is $[Ar]\ 4s^2\ 3d^{10}\ 4p^5$ or $\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}\ \boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow}$

15) see answers on page 574.

16.

a) smallest atomic radius in period 3
b) lowest IE1 in period 5
c) highest EA in **<u>group</u>** 17 (VIIA)- this is the atom that wants an electron the most and will release the greatest amount of energy when it gains an electron (EA will be large and negative)
d) largest atomic radius in period 6
e) most metallic character in **<u>group</u>** 14 (IVA)
h) has condensed ground state electron configuration [Ne] $3s^2\ 3p^2$
i) period 4 element whose atom has a full outer energy level
j) transition metal for which the aufbau principle predicts an electron configuration that would be incorrect



20. You can not measure the size of an atom directly because atoms are not solid spheres. The size of an atom is determined by the distance of the outermost electrons from the nucleus. Because electrons have wave-like behaviour and are constantly being attracted to the nucleus and repelled by other electrons, their position is constantly changing, so it is impossible to measure how far an electron will be from the nucleus.

21. See answers on page 574.

21. Practice multiple choice questions:

a) Answers to multiple choice questions for Chapter 3:

1. a	11. c	21. a	31. a	41. a	51. b	61. d
2. d	12. c	22. d	32. b	42. d	52. d	62. c
3. d	13. d	23. b	33. b	43. b	53. c	63. c
4. d	14. b	24. d	34. c	44. c	54. d	64. d
5. c	15. d	25. c	35. c	45. b	55. c	65. c
6. a	16. a	26. c	36. c	46. a	56. a	66. a
7. a	17. b	27. c	37. d	47. c	57. d	
8. b	18. a	28. c	38. c	48. d	58. a	
9. d	19. b	29. d	39. b	49. a	59. b	
10. d	20. a	30. a	40. b	50. a	60. b	