

NAME: JPV
BLOCK: 101
DATE: _____

ATOMIC STRUCTURE TEST

PART I - Atomic Model

1. In detail discuss the following atomic models (2 marks each)

a) Rutherford Model:

— A dense positively charged nucleus with electrons moving around the nucleus.

b) Dalton Model:

— BILLIARD BALL.

PART II - Atomic Structure

2. How was the electron discovered? What evidence proved that the beam was actually a particle beam and not a beam of light? (2 marks)

→ CROOKER'S TUBE. A magnet deflected the particle BEAM.

3. What is the difference between an isotope and an ion? (2 marks)

Isotope is the same element with a different # of n , while ions have a different # of e^-

4. In terms of atomic structure, what is an atom's atomic number?

the # of p^+

5. What distinguishes an atom of nitrogen from an atom of oxygen?

the # of p^+

6. Complete the following table: (5 marks)

SYMBOL	CHARGE ON ION	NUMBER OF PROTONS	NUMBER OF ELECTRONS	NUMBER OF NEUTRONS
$^{133}_{55}\text{Cs}$	0	55	55	78
$^{115}_{48}\text{Cd}$	2+	48	46	67
$^{36}_{17}\text{Cl}$	1-	17	18	19
$^{40}_{20}\text{Ca}$	2+	20	18	20
$^{199}_{81}\text{Tl}$	0	81	81	118

PART III - The Bohr Model

7. State the number of electrons found in each of the energy levels in the Bohr model. (1 mark)

2, 8, 8
1st 2nd 3rd

8. Draw the Bohr Model of the atom for each atom/ion indicated. (3 marks each)

a) S)))
2e- 8e- 6e-

b) Mg ion $\rightarrow$ Mg^{2+}))
2e- 8e-

9. A general comparison between the electron configurations of ions and noble gases is: they are isoelectronic.

PART IV - The Quantum Mechanical Model

10. Write the ground state electron configuration of chlorine and chromium (Cr) atoms, labelling core electrons and outer electrons. (2 marks)

Cl $1s^2 2s^2 2p^6 3s^2 3p^5$
CORE OUTER

Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
CORE OUTER

Remember group #6
of group #11
Exceptions?

e^- absorb energy AND release it as a specific wavelength of light (i.e., a spectrum).

11. Describe the Bohr model of the atom. ^R What initial evidence was there that Bohr's theory might be correct? Why was his theory eventually discarded? (3 marks)

- positive, dense nucleus that contains protons & neutrons
- e^- move in orbits @ specific energy levels.

this theory was discarded b/c further analysis of spectra revealed that the exact location of e^- was not known

12. How do the 1s and the 2s orbital differ? How are they alike? (2 marks)

different energy
2s is higher

same shape
sphere

13. Using the quantum mechanical model, write the electron configuration for the following (5 marks)

ELEMENT/ION

ELECTRON CONFIGURATION

oxygen

$1s^2 2s^2 2p^4$

fluoride ion (F^-)

$1s^2 2s^2 2p^6$

sodium ion (Na^+)

$1s^2 2s^2 2p^6$

Sn

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$

nitrogen ion (N^{3-})

$1s^2 2s^2 2p^6$

***** BONUS *****

Chemically speaking how does a T.V. work? (2 marks)

e^- hit a screen that is coated with a surface that gives off a flash of light. the colours we see depends on the energy of the e^-

Periodic table key starts on the next page down. Keep scrolling!

Worksheet 4-3

Periodic Trends

Glencoe Chemistry pp.163-169

Name KEYPeriod

ATOMIC RADIUS

1. What trend in atomic radius do you see as you go down a group/family on the periodic table?

INcrease

2. What causes this trend?

more e^- occupy higher energy shells.

3. What trend in atomic radius do you see as you go across a period/row on the periodic table?

*Decrease**from left to right*

4. What causes this trend?

INcrease in nuclear charge AND while the total # of e^- increases the energy shell does NOT so e^- get sucked in a little

5. Circle the atom in each pair that has the largest atomic radius.

a)

Al~~B~~

b)

S~~O~~

c)

Br~~Cl~~

d)

Na

Al

e)

O

F

f)

~~Mg~~Ca

IONIZATION ENERGY

6. Define ionization energy.

the energy required to remove the outermost e^- from a neutral atom.

7. Is it easier to form a positive ion with an element that has a high ionization energy or an element that has a low ionization energy?

Low.

8. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Explain your answer.

 Na^+ b/c there are fewer p^+ to pull on e^-

9. a. Explain why it is harder to remove an inner shell electron than a valence electron from an atom.

b/c there e^- are closer to the nucleus.

- b. Explain why sodium forms a 1+ ion (Na^+) but magnesium forms a 2+ ion (Mg^{2+}).

Na has 1 valence e^- while Mg has 2.

10. What trend in ionization energy do you see as you go down a group/family on the periodic table?

Decrease b/c e^- that are further from the nucleus are easier to remove.

11. What causes this trend?

*See Q #10.*12. What trend in ionization energy do you see as you go across a period/row on the periodic table?*INC.**from left to right.*

13. What causes this trend?

Atomic radius decreases making it more difficult to remove valence e^-

14. Circle the atom in each pair that has the greater ionization energy.

a) Li Beb) Na Kc) Cl Sid) Ca Bae) P Arf) Li K**ELECTRONEGATIVITY**

15. Define electronegativity

the affinity atom has for e^- 16. How does the ionic radius of a nonmetal compare with its atomic radius?*larger than* 

17. What trend in electronegativity do you see as you go down a group/family on the periodic table?

Decrease.

18. What causes this trend?

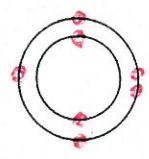
*the larger atomic radius means the pull of pt on e^- will be weaker.*19. What trend in electronegativity do you see as you go across a period/row on the periodic table?*INC.**from left to right.*

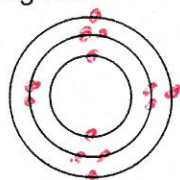
20. What causes this trend?

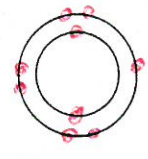
Dec. Atomic radius AND shells that are almost full inc the tendency to attract e^-

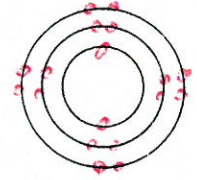
21. Circle the atom in each pair that has the greater electronegativity.

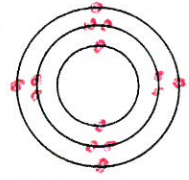
a) Ca Gab) Li Oc) Cl Sd) Br Ase) Ba Srf) O S

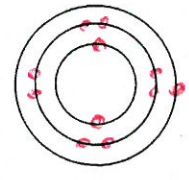
$\frac{7}{\text{N}}$ <i>Nitrogen</i> $\frac{14.0}{}$	P = <u>7</u> N = <u>7</u> E = <u>7</u>
Bohr Diagram	
	
Lewis Structure	N

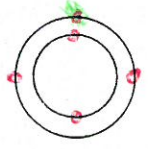
$\frac{13}{\text{Al}}$ <i>Aluminum</i> $\frac{27.0}{}$	P = <u>13</u> N = <u>14</u> E = <u>13</u>
Bohr Diagram	
	
Lewis Structure	Al

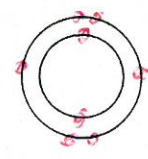
$\frac{9}{\text{F}}$ <i>Fluorine</i> $\frac{19.0}{}$	P = <u>9</u> N = <u>10</u> E = <u>9</u>
Bohr Diagram	
	
Lewis Structure	F

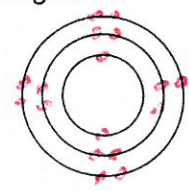
$\frac{18}{\text{Ar}}$ <i>Argon</i> $\frac{39.9}{}$	P = <u>18</u> N = <u>22</u> E = <u>18</u>
Bohr Diagram	
	
Lewis Structure	Ar

$\frac{14}{\text{Si}}$ <i>Silicon</i> $\frac{28.1}{}$	P = <u>14</u> N = <u>14</u> E = <u>14</u>
Bohr Diagram	
	
Lewis Structure	Si

$\frac{11}{\text{Na}}$ <i>Sodium</i> $\frac{23.0}{}$	P = <u>11</u> N = <u>12</u> E = <u>11</u>
Bohr Diagram	
	
Lewis Structure	Na

$\frac{4}{\text{Be}}$ <i>Beryllium</i> $\frac{9.0}{}$	P = <u>4</u> N = <u>5</u> E = <u>4</u>
Bohr Diagram	
	
Lewis Structure	Be

$\frac{8}{\text{O}}$ <i>Oxygen</i> $\frac{16.0}{}$	P = <u>8</u> N = <u>8</u> E = <u>8</u>
Bohr Diagram	
	
Lewis Structure	O

$\frac{17}{\text{Cl}}$ <i>Chlorine</i> $\frac{35.5}{}$	P = <u>17</u> N = <u>19</u> E = <u>17</u>
Bohr Diagram	
	
Lewis Structure	Cl

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