Chemistry 20 – Lesson 2
Atoms, ions, compounds

/100

Part 1

Group 1 / Group IA							18 / VIIIA
1 1e⁻ 1p⁺ hydrogen H	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	2 2e⁻ 2p⁺ helium He
3 1e⁻ 2e⁻ 3p⁺ lithium Li	4 2e⁻ 2e⁻ 4p⁺ beryllium Be	5 3e⁻ 2e⁻ 5p⁺ boron B	6 4e⁻ 2e⁻ 6p⁺ carbon C	7 5e⁻ 2e⁻ 7p⁺ nitrogen N	8 6e⁻ 2e⁻ 8p⁺ oxygen O	9 7e⁻ 2e⁻ 9p⁺ fluorine F	10 8e⁻ 2e⁻ 10p⁺ neon Ne
11 1e⁻ 8e⁻ 2e⁻ 11p⁺ sodium Na	12 2e⁻ 8e⁻ 2e⁻ 12p⁺ magnesium Mg	13 3e⁻ 8e⁻ 2e⁻ 13p⁺ aluminum Al	14 4e⁻ 8e⁻ 2e⁻ 14p⁺ silicon Si	15 5e⁻ 8e⁻ 2e⁻ 15p⁺ phosphorous P	16 6e⁻ 8e⁻ 2e⁻ 16p⁺ sulfur S	17 7e⁻ 8e⁻ 2e⁻ 17p⁺ chlorine Cl	18 8e⁻ 8e⁻ 2e⁻ 18p⁺ argon Ar

Questions:

1. What is the relationship between the old American system group number and the number of valence electrons?

   The roman numeral matches the number of valence electrons.

2. What is the relationship between the period number and the number of energy levels in which electrons are accommodated?

   The period number is the same as the number of electron energy levels for the atoms in the period.

3. What is the relationship between the maximum number of electrons in each energy level and the number of atoms in each period of the periodic table?

   The number of atoms in a period equals the maximum number of electrons that can exist at that energy level.

4. According to the above abbreviated periodic table, how many electrons can be accommodated before a new energy level is started in each of the first three energy levels?

   1st energy level **2** 2nd energy level **8** 3rd energy level **8**

5. Do the diagrams drawn above represent what the electron is actually doing? Explain.

   While the actual motions of the electrons are essentially unknown, the energies of the electrons can be measured. Thus, the energy level diagrams indicate energy only.

25 marks

Part 2

Group 1 Group IA							18 VIIIA
1 $1p^+$ hydrogen H^+	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	
3 ~~$2e^-$~~ $3p^+$ lithium Li^+	4 ~~$2e^-$~~ $4p^+$ berylium Be^{2+}			7 ~~$8e^-$~~ ~~$2e^-$~~ $7p^+$ nitride N^{3-}	8 ~~$8e^-$~~ ~~$2e^-$~~ $8p^+$ oxide O^{2-}	9 ~~$8e^-$~~ ~~$2e^-$~~ $9p^+$ fluoride F^-	
11 ~~$8e^-$~~ ~~$2e^-$~~ $11p^+$ sodium Na^+	12 ~~$8e^-$~~ ~~$2e^-$~~ $12p^+$ magnesium Mg^{2+}	13 ~~$8e^-$~~ ~~$2e^-$~~ $13p^+$ aluminum Al^{3+}		15 ~~$8e^-$~~ ~~$8e^-$~~ ~~$2e^-$~~ $15p^+$ phosphide P^{3-}	16 ~~$8e^-$~~ ~~$8e^-$~~ ~~$2e^-$~~ $16p^+$ sulfide S^{2-}	17 ~~$8e^-$~~ ~~$8e^-$~~ ~~$2e^-$~~ $17p^+$ chloride Cl^-	

Questions:

1. What relationship exists between the electron structure of a Group A ion and the electron structure of the nearest noble gas?
   They have the same electron structure.

2. Why do boron, carbon and silicon not form simple ions? How do they satisfy their electron requirements?
   Boron, carbon and silicon prefer sharing electrons with other atoms to form molecular bonds rather than exchanging electrons to form ionic bonds.

3. What charge do the ions from the following groups assume?

Group 1(IA)	Group (IIA)	Group (IIIA)	Group (VA)	Group (VIA)	Group (VIIA)
1+	**2+**	**3+**	**3–**	**2–**	**1–**

4. What evidence is there that a noble-gas-like electron structure is stable?
   When atoms form ions their electron structure becomes that of the nearest noble gas.

5. What are the differences in the chemical properties of a sodium atom and a sodium ion?

   A sodium atom has one valence electron. In order to lose the electron to have the electron structure of a noble gas (neon) it reacts with a non-metal and loses its electron to the non-metal. Thus a sodium atom will readily react with other atoms. A sodium ion already has a stable electron structure and therefore it does not have to react with something else.

25 marks

Part 3
Part A:

1. Atoms with the same atomic number but a different atomic mass are called **isotopes**.
2. The results of the gold foil experiment led **Ernest Rutherford** to suggest that atoms are mostly empty space, but do contain a "solid" core which is called the **nucleus**.
3. Elements in group 15 have **five** electron(s) in the outer-most energy level.
4. The number of valence electrons equals the number of electrons in the **outer or valence** energy level.
5. The halogens have **seven** valence electron(s).
6. Noble gases do not react at SATP because they have **a stable electron structure**.
7. An atom that has lost or gained electrons is called a(n) **ion**.
8. If an atom gains electrons, it forms **negative** charged ions called **anions**.
9. When electrons are lost from an atom it forms **positively** charged ions called **cations**.
10. Elements that do not chemically react with other elements are said to be **chemically inert**. An example is **helium, neon, argon, krypton, xenon, or radon**.
11. The alkali metals form ions with a **+1** charge.
12. When elements in group 2 react with other elements, they **lose 2** electrons.
13. When elements in group 16 react with other elements, they **gain 2** electrons.
14. Metals tend to **lose** electrons, whereas non-metals tend to **gain** electrons.
15. When there is a transfer of electrons from one atom to another, a(n) **ionic** bond is formed.
16. When two atoms share electrons a **molecular or covalent** compound forms.
17. **Ionic** compounds are formed when a metal reacts with a non-metal.
18. The subatomic particle that is much smaller than the others is the **electron**.
19. When two non-metals are combined they form a(n) **molecular/covalent** compound.
20. When a metal and a non-metal combine they form a(n) **ionic** compound.
21. The noble gas with electrons only in the first energy level. **helium**
22. The halogen that forms ions containing 18 electrons. **chlorine**
23. The element in period 3 containing 3 valence electrons. **aluminum**

23 marks

Part B: Complete the chart below.

Atom or ion name	Symbol	Atomic mass	Atomic number	# of neutrons	# of protons	# of electrons	# of electrons lost/gained	Net charge
iron (III)	Fe^{3+}	56	26	30	26	23	lost 3	3+
manganese (IV)	Mn^{4+}	55	25	30	25	21	lost 4	4+
sodium	Na^{+}	23	11	12	11	10	lost 1	1+
aluminum	Al^{3+}	27	13	14	13	10	lost 3	3+
argon	Ar	4040	18	22	18	18	0	0
fluoride	F^{-}	19	9	10	9	10	gained 1	1–
hydride	H^{-}	1	1	0	1	2	gained 1	1–
sulfide	S^{2-}	32	16	16	16	18	gained 2	2–
magnesium	Mg^{2+}	24	16	12	12	10	lost 2	2+
calcium	Ca^{2+}	40	20	20	20	18	lost 2	2+
hydrogen	H^{+}	1	1	0	1	0	lost 1	1+
sulfur	S	32	16	16	16	16	0	0
cobalt (III)	Co^{3+}	59	27	32	27	24	lost 3	3+

12 marks

Part C: Complete the following table. Note the name of a *non-metallic* ion ends in *ide* while the name for a *metallic* ion uses the full name of the metal.

Ion name	**Symbol**	**# of p^+**	**# of e^-**	**# of electrons lost/gained**	**Same electron structure as what Noble gas?**
e.g. fluoride	F^-	9	10	gained 1	neon
1. iodide	I^-	53	54	gained 1	xenon
2. sulfide	S^{2-}	16	18	gained 2	argon
3. potassium	K^+	19	18	lost 1	argon
4. calcium	Ca^{2+}	20	18	lost 2	argon
5. bromide	Br^-	35	36	gained 1	krypton
6. strontium	Sr^{2+}	38	36	lost 2	krypton
7. hydrogen	H^+	1	0	lost 1	(none)
8. oxide	O^{2-}	8	10	gained 2	neon
9. magnesium	Mg^{2+}	12	10	lost 2	neon
10. aluminum	Al^{3+}	13	10	lost 3	neon
11. selenide	Se^{2-}	34	36	gained 2	krypton
12. hydride	H^-	1	2	gained 1	helium
13. lithium	Li^+	3	2	lost 1	helium
14. rubidium	Rb^+	37	36	lost 1	krypton
15. chloride	Cl^-	17	18	gained 1	argon

15 marks