Unit 1, Lesson 01: Summary of Atomic Structure so far…

Atoms are made of sub-atomic particles:

- Protons: found in the **nucleus**, charge of **1+**, mass of *1 amu* (u)
- Neutrons: found in **nucleus**, **no** charge, mass of **1 amu** (u)
- Electrons: found in **the space** around nucleus, charge of **1 –**, mass is **1/1837 amu (negligible)**

Atomic Number: the number of **protons** in the nucleus of an atom

- symbol is " **Z** "
- defines the **identity** of the atom

eg. Z = 12, atom is **magnesium (Mg)**   Z = 47, atom is **silver (Ag)**

Mass Number: the number of **protons** + **neutrons** in the nucleus of an atom

- symbol is " **A** "
- it is a **counted** value, it has **no units**
- it is **NOT** reported on periodic table
- isotopes are identified by their mass number

mass number

Examples:

Isotope	**Pb – 206**	**Pb – 207**	**Pb - 208**
Atomic Number (Z)	**82**	**82**	**82**
Mass Number (A)	**206**	**207**	**208**
Number of Neutrons (A – Z)	**124**	**125**	**126**

Average Atomic Mass (AAM): the **weighted** average mass of all **isotopes** of an element

- reported on periodic table, units are amu (u)
- AAM on periodic table is usually close to the mass number of the most **abundant** (**common**) isotope

eg. AAM of carbon is **12.01 amu** so most abundant isotope is probably **C – 12**

eg. most abundant isotope of argon is probably **Ar – 40**

Ions: charged atoms

- if the number of electrons equals the number of protons, the atom is **neutral** (**uncharged**)
- if there are more electrons than protons, the ion is **negatively** charged (an **anion**)
- if there are fewer electrons than protons, the ion is **positively** charged (a **cation**)
- atoms tend to gain or lose electrons to obtain a **stable octet** electron arrangement and become **isoelectronic** with a **Noble Gas** (have the same number and arrangement of electrons as a Noble Gas)

eg. Mg loses **2** electrons, forms **Mg^{2+}** ions

Standard format:

$^{16}_{8}O^{2-}$

mass number (A)
charge
chemical symbol
atomic number (z)

protons (atomic number, Z) = **8**
neutrons (mass number – atomic number) = **8**
electrons (2 more electrons than protons) = **10**

Unit 1, Lesson 01: Answers to Homework

Textbook Questions Page 130: 1, 2, 4, 5, 6, 7

1. In Thomson's "Raisin Bun" model:

- the atom is a solid sphere
- the solid part of the atom is positively charged, called a matrix
- the electrons are stuck in the matrix

In Rutherford's "Electron Cloud" model:

- the atom is mostly empty space
- the centre of the atom is positively charged and dense, called the nucleus
- the electrons are flying randomly around the nucleus in a cloud

Rutherford's model was based on the gold foil experiment. He fired positively charged alpha particles (helium nuclei) at a piece of gold foil that was only a few atoms thick.

- Most of the alpha particles passed directly through the gold foil, which indicated that atoms are not solid particles. Rather, atoms are mostly empty space.
- A tiny number of alpha particles were deflected from the gold foil, which indicated that there was a strong positively charged region in the atom.
- A very tiny number of alpha particles bounced right back, which indicated that atoms contain a small, very dense positively charged "core" that the alpha particles bounced off of. Rutherford called this the nucleus.

2. Rutherford believed that the electrons were found in a random cloud around the nucleus. If this were true, then all atoms would give off an identical continuous spectrum that included all colours of light when they were excited. However, when Bohr studied the emission spectrum of hydrogen, he saw only certain colours of light. Each colour represented electrons with a certain energy. Bohr realized that because there were only certain colours of light, that the electrons were always certain specific distances from the nucleus. He suggested that electrons moved around the nucleus in defined, predictable orbits that had certain discrete energy levels called "quantum levels".

4. The types of electromagnetic radiation from:

a) longest to shortest wavelength (ie. from lowest to highest energy) are:

radio waves → microwaves → infrared (heat) → visible light → ultraviolet light → X-rays → gamma rays

b) highest to lowest frequency (as wavelength decreases, frequency increases):

gamma rays → X-rays → ultraviolet light → visible light→ infrared (heat) → microwaves → radio waves

5. Compare and contrast visible light and ultraviolet light:

visible light	**ultraviolet light**
• type of electromagnetic radiation • speed in a vacuum is 3.00 x 10^8 m/s (text has a mistake here, should not be 10^{-8}) • wavelength : 400 – 750 nm • lower energy • can be seen by the human eye	• type of electromagnetic radiation • speed in a vacuum is 3.00 x 10^8 m/s • wavelength : about 200 – 400 nm • higher energy • can not be seen by the human eye

6a) When an electron moves from n = 1 to n = 6:
- it moves further from the nucleus
- it absorbs energy
- it does not give off light

6b) When an electron moves from n = 5 to n = 2:
- it moves closer to the nucleus
- it will release (give off) energy
- it will give off light of a specific wavelength (energy), the colour will depend on which atom the electron belongs to
- it is a "big jump" so the light will have a short wavelength and be high energy

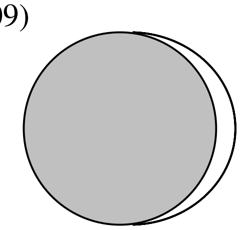
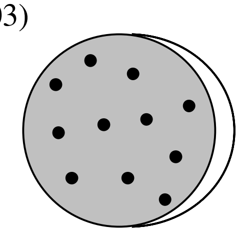
6c) When an electron moves from n = 4 to n = 3:
- it moves closer to the nucleus
- it will release (give off) energy
- it will give off light of a specific wavelength (energy), the colour will depend on which atom the electron belongs to
- it is a "little jump" so the light will have a long wavelength and be low energy

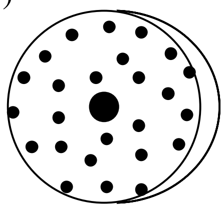
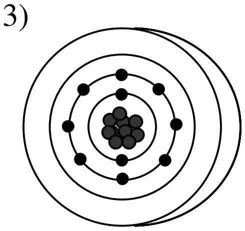
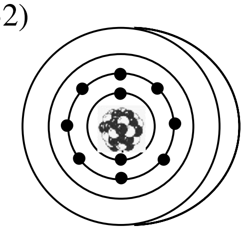
7. From lowest to highest energy, the principal quantum levels are: n = 1, n = 2, n = 4, n = 5 and n = 7
- the larger the principal quantum number, the further the electrons are from the nucleus and the higher their potential energy

Question 3: Explain why Bohr's model of the atom had to be modified. (It has two fundamental flaws).

Bohr's model of the atom (the planetary model) had to be modified because:
- it did not include neutrons
- it could not account the line spectra for atoms or ions with 2 or more electrons

Worksheet Questions
4. Atomic Models from Dalton to Chadwick

Researcher and Model of Atom	**Features and Limitations of this Model of the Atom (include how it is different from previous models)**
John Dalton (1809) 	**Billiard Ball Model: no electrons** • first "modern" view of atoms after the Four Element Theory (Air, Fire, Earth and Water) • atom is seen as a solid piece of indivisible matter • atoms can bond together in definite proportions to form compounds Problem: it couldn't explain Thomson's work with cathode rays, also did not account for isotopes
J. J. Thomson (1903) 	**Raisin Bun Model: electrons embedded in solid, positive matrix** • atom is made of positive matrix and negative particles called electrons • the atom is divisible • the model explained cathode rays as streams of electrons • electrons are a fundamental particle of all matter Problem: it couldn't explain Rutherford's work with gold foil

Ernest Rutherford (1911) 	**Electron Cloud Model: electrons in random cloud around nucleus** • atom is mostly empty space, this is where the electrons are found • centre of atom is a solid, very dense, massive, positively charged region called the nucleus • electrons move freely and randomly around nucleus Problems: • couldn't explain why the nucleus doesn't fly apart because of the repulsion between the positive charges • couldn't explain why the electrons don't just stick to the nucleus • couldn't explain why atoms only give off certain colours of light instead of all colours (Bohr's work) • (also could not explain black body radiation and the photoelectric effect- discussed at university)
Neils Bohr (1913) 	**Planetary Model: electrons in specific energy levels around nucleus** • electrons are not in a random cloud around the nucleus, they are in specific energy levels called principal quantum levels (n) • all electrons in a certain quantum level are exactly the same distance from the nucleus, so they have exactly the same amount of energy • when electrons are excited, they absorb energy and move further from the nucleus • when electrons fall back closer to the nucleus, they release energy of a certain wavelength (colour). The colour depends on the energy difference between the two quantum levels it moves between • explains the bright line spectra of hydrogen Problems: does not explain the missing mass in the nucleus • does not accurately predict spectral lines for atoms with 2 or more electrons
James Chadwick (1932) 	**Planetary Model with Neutrons** • model is same as Bohr's planetary model with regard to electrons • discovery of neutrons accounts for the extra, uncharged mass in the nucleus (the charge:mass ratio) • neutrons account for differences in nuclear stability of atoms of the same element (radio-isotopes) Problem: still does not explain the spectral lines for atoms with 2 or more electrons

Worksheet Questions:

5. Complete the following chart:

Element	Atomic Number (Z)	Number of Protons	Number of Electrons	Overall Charge	Number of Neutrons	Mass Number (A)	Average Atomic Mass (u)	Isotope (eg. Ag – 107)
Mg	12	12	10	2+	12	24	24.31	Mg-24
Ag	47	47	46	1+	62	109	107.87	Ag-109
S	16	16	18	2-	17	33	32.07	S-33
Mg	12	12	12	0	13	25	24.31	Mg-25
Se	34	34	36	2-	46	80	78.96	Se-80
Au	79	79	78	1+	118	197	196.97	Au-197
Mg	12	12	12	0	14	26	24.31	Mg-26
Sr	38	38	36	2+	50	88	87.62	Sr-88
O	8	8	10	2-	8	16	16.00	O-16
Sn	50	50	46	4+	69	119	118.71	Sn-119
Sb	51	51	48	3+	67	118	121.76	Sb-118
Au	79	79	76	3+	117	196	196.97	Au-196
Sb	51	51	46	5+	67	118	121.76	Sb-118
Pt	78	78	77	1+	117	195	195.08	Pt-195
Sr	38	38	38	0	52	90	87.62	Sr-90

6. Referring to the chart above, and using the format "Ag-107", write:

a) any atoms that are isotopes of each other: Mg-24 , Mg-25 and Mg-26; Au-197 and Au-196; and Sr-88 and Sr-90. Note: because both Sb atoms have the same number of neutrons, they are not isotopes of each other, but they are different ions

b) any atoms that are ions of each other: Mg^{0} and Mg^{2+}; Au^{1+} and Au^{3+}; Sb^{3+} and Sb^{5+}; Sr^{0} and Sr^{2+}

7. Using your Periodic Table and the format "Ag-107", predict the most common isotope for each of the following elements:

a) Mg-24
b) Sr-88
c) Al-27
d) H-1
e) S-32
f) Na-23
g) F-19
h) Ar-40
i) Pb-207

8. Using only your Periodic Table, write the ion that will be formed (if any) by each of the following elements:

a) Mg^{2+}
b) S^{2-}
c) Al^{3+}
d) Br^{1-}
e) N^{3-}
f) Ba^{2+}
g) Cs^{1+}
h) Ne will not ionize
i) Te^{2-}

9. Atoms and ions that have the same number of electrons are said to be **isoelectronic**. List three atoms or ions, including the charge on each, that are isoelectronic with the following:

a) Ar: Cl^{1-}, S^{2-}, P^{3-}, K^{1+}, Ca^{2+}
b) Br^{1-}: Kr, Se^{2-}, As^{3-}, Rb^{1+}, Sr^{2+}
c) Ne: F^{1-}, O^{2-}, N^{3-}, Na^{1+}, Mg^{2+}
d) Sr^{2+}: Kr, Br^{1-}, Se^{2-}, As^{3-}, Rb^{1+}
e) S^{2-}: Ar, Cl^{1-}, P^{3-}, K^{1+}, Ca^{2+}
f) Na^{1+}: Ne, F^{1-}, O^{2-}, N^{3-}, Mg^{2+}