Unit 1 Review: Matter and Chemical Bonding

1. Do you think DHMO should be banned? Justify your answer. Write the formula for dihydrogen monoxide.
   H_2O
2. Name these groups on the periodic table: 1, 2, 17, 18, 3-12, and the 2 rows at the bottom of the table.
   1 – Alkali metals; 2 – Alkali Earth metals, 17 – Halogens, 18 – Nobel gas; 3-12 – Transition elements; 2 rows at the bottom – Lanthanoids and Actinoids.
3. Summarize the contributions of Dalton, Thompson, Rutherford, Bohr, and Chadwick.
   Dalton – The atomic theory – all matter is made up of atoms, atoms are like spheres
   Thompson – negatively charged particles embedded in a positive atom.
   Rutherford – Most of the atom is empty space; dense, centrally located nucleus that is positively charged; orbiting around are negatively charged electrons
   Bohr: Electrons are in specific energy levels orbiting around the nucleus
   Chadwick – found neutrons
4. Who developed the first periodic table? How are the elements ordered?
   Mendeleev; Elements are arranged by increasing atomic mass.
5. Define atomic number and atomic mass. Why are atomic masses not whole numbers?
   Atomic number – The number of protons in the nucleus of an atom, which determines the chemical properties of an element and its place in the periodic table
   Atomic mass – The mass of an atom of an element expressed in atomic mass units. It is approximately equivalent to the number of protons and neutrons in the atom
   Why atomic mass not whole numbers – Because there are different isotopes of each element. The atomic mass is the average based on the proportions of all the isotopes that occur naturally.
6. A neutral atom of Cl-37 has **20** neutrons. Its mass # is **37,** its atomic # is **17**, and it has **17** electrons.
7. How does atomic size change going down a group or across a period? Explain.
   Going down a group – number of energy levels increase – therefore the radial distance and electron shielding increases – the atomic radius increases.
   Going across a period – Same number of energy levels therefore no effect of radial distance and electron shielding. The number of protons in the nucleus increases therefore the nuclear charge increases exerting more pull on its valence electrons. This shrinks the atomic size as you go from left to right of the periodic table.
8. Define ionization energy, electron affinity, and electronegativity. How do these change as we move down a group or across a period? Explain each trend (referring to atomic size, nuclear charge, etc.).
   Ionization energy – energy required to remove an electron.
   T→ B: Atomic size increases because of more number of energy levels and more electron shielding therefore the electron is farther away from the nucleus – less energy required to remove the electron
   L→R: Atomic radius becomes smaller because more pull on the electrons due to more number of protons in the nucleus. Therefore removing an electron becomes more difficult or more energy required to remove an electron.
   Electron affinity – energy released or absorbed when an electron is added to a neutral atom
   T→ B: Decreases as less energy is released when an electron is added because the size of the atom becomes larger and the pull on electrons by the protons in the nucleus is weaker
   L→ R: Increases as the atomic size is smaller so the nuclear pull on electrons stronger so more energy released when an electron is added
   Electronegativity – is the ability of an atom to attract shared electrons
   T→ B: Decreases as the number of energy shells increase the nuclear pull on the electrons in the valence orbit is weaker.
   L→ R: Increases as the size decreases the nuclear pull on the electrons is more stronger

9. Draw Lewis diagrams for: O, Al, Na, I, Xe. Draw the Lewis structures for their ion?

10. Identify each as ionic or covalent: a) CO_2, b) NaCl, c) $FeCl_2$, d) CCl4, e) Al_2SO_3.
    CO_2 – Molecular ; NaCl – Ionic; $FeCl_2$ – Ionic; CCl_4 – Molecular; Al_2SO_4 – Ionic.
11. Illustrate the bonding between Mg and P using the three steps (equation) to show it's formation.

12. Use Lewis diagrams to show the bonding between O + Cl, P + H, Ca + Cl, N + N.

13. List the typical physical properties of ionic & covalent compounds? What causes these differences?
    Ionic compounds are solids in room temperature, soluble in water, conductors of electricity when melted or in solution with high melting and boiling points.
    Molecular compounds are solids, liquids or gases at room temperature that are non conductors of electricity with low melting and boiling points.
    Ionic bonds are stronger than covalent bonds thus requires lots of energy to break them thus have high melting and boiling points.
    When dissolved or melted the ions in ionic compounds move freely thus allowing the free flow of electrons.
14. Classify each bond as ionic, covalent, slightly polar covalent, or polar covalent: H_2, CH_4, LiF, H_2O.
    H_2 – pure covalent ΔEN = 0; CH_4 – pure covalent ΔEN = 2.6 -2.2 = 0.4 ; LiF – Ionic ΔEN = 4-1 =3; H_2O – polar covalent ΔEN = 3.4-2.2 = 1.2;
15. Use valence (crossover) to determine the chemical formulae for H + Cl, O + Na, P + Cl, Al + O, Mg + O.
    HCl; Na_2O; PCl_3; Al_2O_3; MgO
16. What are the four molecular shapes? Give an example for each.

Linear	Bent	Pyramidal	Tetrahedral
CO_2	H_2O	NH_3	CCl_4

17. What is the difference between a polar molecule and a non-polar molecule?
    A polar molecule is one in which the bond dipoles present do not cancel each other out and thus results in a molecular dipole (slight negative charge on one end of the molecule and a slight positive charge on the other end of the molecule). Cancellation depends on the shape of the molecule or Stereochemistry and the orientation of the polar bonds.
    Non-polar molecules do not have molecular dipoles.
18. Draw the following molecules and indicate if they are polar or nonpolar: a) SeH_2, b) CH_3Cl c) HBr

    SeH_2 has slightly polar bonds (ΔEN =0.4)and it's shape is bent → polar molecule
    CH_3Cl: C-H bonds are slightly polar (ΔEN =0.4) and C-Cl are polar (ΔEN =0.6), the shape is tetrahedral, due to unequal dipoles the molecule is polar
    HBr: H-Br bond is polar (ΔEN =0.8) and shape is linear. The molecule is polar

19. What is the difference between intermolecular forces and intramolecular forces?
    <u>Intra-molecular forces are the force within the molecule – Ionic, covalent, coordinate covalent and metallic bond</u>
    <u>Inter-molecular forces are forces of attraction between molecules – London dispersion forces, Dipole – dipole forces, Hydrogen bonds</u>
20. What are the three types of intermolecular forces?
    <u>London dispersion forces – weakest force seen in all molecules</u>
    <u>Dipole-dipole force – seen between molecules that are polar</u>
    <u>Hydrogen bonds – strongest of all seen when the molecule has hydrogen and oxygen or nitrogen or fluorine</u>
21. What forces would be involved in the following molecules: a) NH_3 b) HBr Who would have the higher mp?
    <u>NH_3: Hydrogen baond and London dispersion forces</u>
    <u>HBr: Dipole dipole forces and London dispersion forces.</u>
    <u>NH_3 would have the higher melting point as more energy is required to break the hydrogen bonds in NH_3 than the dipole dipole forces in HBr.</u>

Additional review questions can be found in the textbook: Unit 1 Self-Assessment Q# 1-8, 11-20, 23, 24, 25 pg. 102-103 Unit 1 Review Q# 1-24, 26-30, 36-40, 42-45, 47, 53-55, 57-61, 66 pg. 96-101

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1 - A; 2 – B; 3 – A; 4 – C; 5 – D; 6 – B; 7 – D; 8 – D; 9 – B; 10 – A;

11 – Group 15, Period 3 based on valence electrons;

12 - An ISOTOPE is an atom of an element with a different number of neutrons than the original element RADIOISOTOPES, however, are unstable by nature due to several things going on in the nucleus. They are isotopes by definition, but they have unusually high levels of unstable energy in the nucleus and thus emit this energy in the form of radiation.

13 – Average atomic mass = (106.91x51.84/100) + (108.90 x 48.16/100) = 55.422144 + 52.4424 = 107.868384

14 – Both are gases; have only one energy level. Helium has a full shell so stable and un-reactive. Hydrogen on the other hand has one valence electron and needs one more to become stable so very reactive.

15 – Atomic radius decreases as you go across the period while electronegativity increases as you go across the period. Because the radius decreases and the nuclear charge increases the pull on the shared electron increases.

16 –

17 – As you go down the group the number of energy levels increase thus increasing the atomic radius and energy shielding. Therefore the valence electron is going to be farther away from the nucleus as you go down the group. Thus removing this electron require less energy than an electron which is closer to the nucleus.

18 – a) Barium +2 (loses two electrons to become stable); b) Sulfur -2 (accepts two electrons to become stable); c) Potassium +1 (loses one electron to become stable); d) Nitrogen -3 (accepts three electrons to become stable)

19 – a) Mg_3N_2 – ionic; b) OF_2 – molecular; c) $SnBr_2$ – ionic; d) $AlPO_4$ – ionic; e) $Co_2(SO_3)_3$ – ionic

20 – a) phosphorous pentachloride – molecular; b) lithium carbonate – ionic; c) copper (II) oxide – ionic; d) dinitrogen trioxide – molecular; e) ammonium nitrite – ionic.

23 – Ionic → high MP, conducts electricity when melted or dissolved in water, solid at room temperature and soluble in water.

– Molecular → low MP, do not conduct electricity, solid, liquid or gas at room temperature and generally not soluble in water.

24 – Salt when dissolved in water will conduct electricity while sugar dissolved in water will not; The melting point of salt is much higher than sugar.

25 – a)

b) The C-H is slightly polar **((ΔEN =0.4) while that between C-Cl is polar (ΔEN =0.6) That is the chlorine atom is going to pull the electrons more towards itself than the carbon. This leads to unequal pull making the molecule polar, even though both molecules have a tetrahedral shape.**
c) Because CH_3Cl is a polar molecule it has DDF and LDF acting between the molecules while CH_4 only has LDF which is a weak force of attraction. Thus the mp and bp of CH_3Cl is much higher than CH_4 as more energy is required to break the bonds between the molecules.

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1 – A; 2 – C ; 3 – D ; 4 – D ; 5 – C ; 6 – A ; 7 – E ; 8 – E ; 9 – B ; 10 – C ;
11 – Thomson – negatively charged particles called electrons scattered in a positively charged atom.

- Rutherford – Most of the atom is empty space; centrally located dense positively charged nucleus; electrons orbiting around the nucleus.

12 – Valence electrons are the electrons in the outermost orbit. These electrons are lost to form a cation or few more electrons are gained to form an anion or are shared to get a stable octect. These affect the properties of the compound.
13 – Isotopes have more neutrons but because their number of protons is still the same does not affect its properties.
14 – Metals are conductors of electricity while non-metals are not.
15 – C<O<Ge<Ca<Ba
16 – K<Ca<Mg<P<Cl
17 – Increases as more protons in the nucleus are going to pull the shared electrons more towards itself.
18 – O_2 (share two pairs of electrons), O^{2-} (gain two electrons), O_3 (share three pairs of electrons with two other oxygen molecules)
19 – Single bond – share one pair of electrons, double bond – share two pairs of electrons, triple bond – share three pair of electrons.
20 – Between Na and N (ΔEN =2.1) → ionic bond; Between S and O (ΔEN =0.8) → polar covalent; between F and F (ΔEN =0) → pure covalent
21 – A metal and a non-metal combine to form an ionic compound.
For eg. NaCl sodium chloride where sodium is a metal and chlorine is a non-metal.
CaO – calcium oxide where calcium is a metal and oxygen is a non-metal
Zn_3N_2 – zinc nitride where zinc is the metal and nitrogen is a non-metal.
22 – A is an ionic compound
B non-polar compound
C polar compound
23 – Metalloids are elements beside the staircase. Their properties are a combination of those of metals and non-metals.
24 – Stable octect means the valence shell is full.
26 – Metals lose electrons to get a stable octect while non-metals accept electrons to get a full valence shell, in the process they become charged.
27 – If there are more intermolecular forces then you require more energy so the physical state will be a solid at room temperature.
28 – When in liquid state the ions in the compound are free to move thus conducting electricity.
29 -

Isotope	Atomic #	Mass number	# of protons	# of neutrons
$^{44}_{20}Ca$	20	44	20	24
$^{20}_{10}Ne$	10	20	10	10
$^{14}_{6}C$	6	14	6	8
$^{37}_{17}Cl$	17	37	17	20
$^{28}_{12}Mg$	12	28	12	16
$^{66}_{30}Zn$	30	66	30	36
$^{138}_{56}Ba$	56	138	56	82

30 – Atomic radius increases as you go down a group; decreases as you go across a period

– Electronegativity decreases as you go down a group and increases as you go across a period.

36 – a) ionic compound is formed because one element loses electrons and another element accepts electron

b) Mg has 2 and Cl has 7

c) Mg has a full second orbit. And Chlorine has 3 full shells.

37 – Ionic compounds are soluble in water because both have charges. Like dissolves like. While when it comes to non-polar compounds they are not polar and do not dissolve in polar solvents.

38 – NaCl would have the higher MP as it is an ionic compound. More energy is required to break up the crystal lattice in NaCl than the intermolecular forces between BrCl molecules.

39 – 2+; 2- and overall charge will be zero; 1- and the overall charge would still be zero.

40 – a) the first diagram is testing a solid ionic compound and the second diagram is testing an ionic compound in solution.

b) No conductivity in the first diagram while the second one is conducting electricity

c) In the first diagram the ions are held due to the crystal lattice while in the second diagram the ions are able to move thus conducting electricity.