Robyn

Science 10: Course Review

Unit 1: Chemistry

1. The following scientists are associated with models of the atom. Describe their model.
   a. Rutherford ~ 'Solar System model' → to account for the deflection of alpha (+) particles. Centre of every atom has a positively charged nucleus. Nucleus accounts for majority of mass, electrons orbit around nucleus, in solar system like pattern. (neutron = no charge, but similar mass to a proton.
   b. Dalton ~ 'Billiard Ball model' → to explain the interaction of chemicals. Matter is composed of small, indivisible particles (atoms), atoms of the same element are identical in mass & size. Atoms are in constant motion & reactions change the way that atoms are arranged.

   * c. Bohr ~ "bullseye" / energy levels ~ electrons orbit specific orbitals called energy levels / electron cells, transitions of electrons to higher energy levels requires energy, and transitions to lower energy levels produce electromagnetic radiation.

   d. Thomson ~ 'Raisin Bun Model' → to explain negative charges in the atom. Atom = positively charged sphere where negatively charged electrons are embedded.
   - Atom has no charge

2. Complete the following table

Subatomic Particle	Relative Charge	Symbol	Location	Mass (g)
proton	positive (+1)	p^+	nucleus of atom	1.67×10^{-24}
neutron	no charge (0)	n^0	nucleus of atom	1.67×10^{-24}
electron	negative (-1)	e^-	outer region of atom	9.02×10^{-28}

3. What is the atomic number? How is it related to number of protons? Where is it located on periodic table? The atomic # is the number of protons in the nucleus of an atom of a particular element. It is located on the top left hand corner of each element, on the periodic table.

4. What is atomic mass? Where is it located on periodic table?
   Atomic mass is the mass of one mol of atoms of an element, expressed in g/mol and is located in the top right corner of each element. = p+n = (19)

5. What is an isotope? Explain using an example.

An isotope is an atom that has the same number of protons, but different numbers of neutrons; therefore having a different mass. Ex ~ carbon 12 & carbon 13 ~but are chemically alike.

6. Using nuclear notation, indicate the atomic number, mass number (of the most common isotope) and symbol of the following atoms.

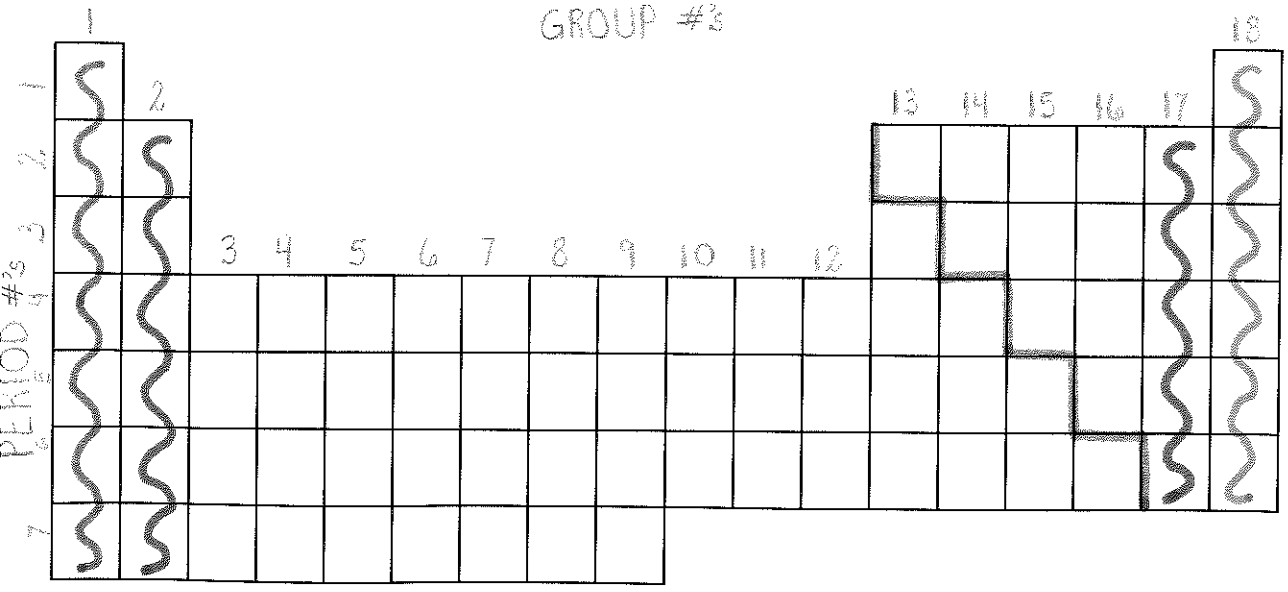
Fluorine 19.00 1- 9 F	**Sodium** 22.99 1+ 11 Na	**Argon** 39.95 18 Ar	**Magnesium** 24.31 2+ 12 Mg

7. Using nuclear notation, indicate the atomic number, mass number and symbol of the following isotopes.

Carbon-12 12 6 C	**Carbon-13** 13 6 C	**Carbon-14** 14 6 C

8. Complete the following tables and then place the labels on the periodic table.

Label	Definition / Characteristic
Period	a horizontal row in the periodic table
Group	a vertical column in the periodic table.
alkali metals	consists of elements in group 1
alkaline earth metals	consists of elements in group 2
halogens	consists of elements in group 17
noble gases	consists of elements in group 18
staircase	



9. Use your periodic table to complete the following:

	Element Name	IUPAC Symbol	Atomic Number	Group Number	Period Number	Metal (m) or Nonmetal(nm)	Family/Series Name
1.	chlorine	Cl	17	17	3	(nm)	halogens
2.	magnesium	Mg	12	2	3	(m)	alkaline earth metals
* 3.	zinc	Zn	30	12	4	(m)	transition metals
4.	nitrogen	N	7	15	2	(nm)	----------
5.	iodine	I	53	17	5	(nm)	halogens
* 6.	gold	Au	79	11	6	(m)	transition metals
7.	Sodium	Na	11	1	3	(m)	alkali metals
* 8.	thorium	Th	90	3	7	(m)	actinoids

10. Fill in the following prefixes on the table below.

Number	Prefix	Number	Prefix
1	mono-	6	hecto-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

11. Summarize the three rules for naming and writing the formula for binary molecular compounds.

#1 ~ use the name of the element farthest to the left on the periodic table.
#2 ~ use the name of the element farthest to the right on the periodic table. (attach -ide to end).
#3 ~ attach a 'prefix' to each name (# of atoms per molecule). → no 'mono' for 1st element.

12. Write the correct chemical name for each of the following molecular compounds:

	Chemical Formula	Chemical Name
1.	$O_{2\,(g)}$	oxygen
2.	$P_2O_{5(s)}$	diphosphorous pentaoxide (solid)
3.	$HCl_{\,(g)}$	hydrogen chloride
4.	$NH_{3(g)}$	ammonia (gas)
5.	$N_2Cl_{4\,(l)}$	dinitrogen tetrachloride (liquid)
6.	$ICl_{5(s)}$	iodine pentachloride (solid)
7.	$CH_{4\,(g)}$	methane

13. What is an ion?

An ion is an atom or group of atoms that carry a (+) or (-) electrical charge.

14. In which states/form do you find ions? (solutions or solids or gases etc)

Solids (s), & solutions (aq).

15. What type of compounds have ions (ionic, molecular, acid,base?) why? Give and example

Ionic, acids, bases.

16. Complete the following table

	English Name	International Symbol	Number of Protons	Number of Electrons	Number of electrons gained	Net Charge
1.	neon atom	Ne	10	10	none	0
2.	lithium ion	Li	3	2	lost 1	1+
3.	Silver ion	Ag	47	46	lost 1	1+
4.	Sulfur ion	S	16	18	gained 2	2-
5.	Silicon atom	Si	14	14	0	0
6.	arsenic ion	As	33	36	gained 3	3-
7.	Cesium ion	Cs	55	54	lost 1	1+
8.	Zinc ion	Zn	30	28	lost 2	2+

17. State the rules for writing the names of binary ionic compounds, and give an example

First, use the full name of the cation (usually metal). Then put the name of the anion (usually non-metal), & add -ide to the ending. NO PREFIXS for ionic compounds.

18. Write the correct chemical name for each of the following ionic compounds:

	Chemical Formula	Chemical Name
1.	$SrCl_{2(s)}$	Strontium chloride
2.	$RbBr_{(s)}$	rubidium bromide
3.	$Na_2O_{(s)}$	Sodium oxide
4.	Al_2S_3 (s)	aluminum sulfide
5.	$ZnCl_2$ (aq)	zinc chloride
6.	$MgCl_2$ (aq)	magnesium chloride

7.	$CoCl_{2(s)}$	Cobalt (II) chloride
8.	$TiO_{2(s)}$	titanium (IV) oxide
9.	$Cu_2O_{(s)}$	Copper (I) oxide
10.	$SnS_{(s)}$	tin(II) sulfide

19. What is a polyatomic ion? Give an example.

a polyatomic ion consists of 2 or more different atoms joined together by covalent bonds, either containing a (+) or (-) charge.

Ex = carbonate (CO_3^{2-})

20. Complete the following chart.

Chemical Formula		Name of Compound
1.	$AlCl_{3(s)}$	Aluminum chloride
2.	$NaI_{(s)}$	Sodium iodide
3.	$MgO_{(s)}$	magnesium oxide
4.	$K_2S_{(s)}$	potassium sulfide
5.	$CaF_{2(s)}$	Calcium fluoride
6.	$GaBr_{3(s)}$	gallium bromide
7.	$MgCO_{3(s)}$	magnesium carbonate
8.	$Na_2SO_{4(s)}$	Sodium sulfate

21. Identify the properties of ionic and molecular compounds

Ionic	Molecular
state at room temp: "Solid" (hard, brittle, crystal lattice shape).	~ solid, liquid, or gas!
melting point = high!	~ low
attraction between molecules: strong crystal lattice.	~ weaker
Conductivity when solid? ~ No, molecules held rigidly in crystal lattice shape.	~ NO

Conductivity when dissolved in water or melted?

~ Yes, electrolyte (ions form)

tend to be non-electrolytes.

25. Fill in the following chart

	Acids	Bases
pH	less than 7. (<7)	greater than 7. (>7)
Litmus	turns blue litmus red.	turns red litmus blue.
bromthymal blue	~yellow~	~blue~
phenolphthalein	~colourless~	~pink~/purple
conductivity	conductive when in solution	conducts electricity ✓
reaction with metal	yes, to produce hydrogen gas.	don't react with zinc/yes
Solubility	yes, to some degree.	some soluble in water (slightly soluble)
Taste	sour!	bitter!

26. Complete the following chart

	ionic (i) molecular (m) or acidic (a)	Chemical Formula	Solubility (high or low)	Chemical Name
1.	ionic	$PbI_{2(s)}$	low	lead (II) iodide
2.	acidic	$HMnO_4$ (aq)	High	permanganic acid
3.	ionic	$NaHS_{(s)}$	----------	Sodium hydrogen sulfide
4.	acidic	H_2SO_3 (aq)	High	sulfurous acid
5.	molecular	$H_2O_{2(l)}$	----------	hydrogen peroxide
6.	ionic	Ti_2O_4 / $TiO_{2(s)}$	High	titanium (IV) oxide
7.	acidic	$HCl_{(aq)}$	High	hydrochloric acid
8.	acidic	$H_2S_{(aq)}$	High	hydrosulfuric acid
9.	ionic	$Ga_2S_{3(s)}$	low	gallium sulphide
10.	acidic	H_2SO_4 (aq)	High	sulfuric acid

27. Write balanced chemical equations for the following reactions (include the states of matter)

water → hydrogen + oxygen

$2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}$

Decom.

nitrogen + hydrogen → ammonia

$N_{2(g)} + 3H_{2(g)} = 2NH_{3(g)}$

Syn.

sulfuric acid + aqueous sodium hydroxide → water + aqueous sodium sulphate

H_2SO_4 (aq) + $2NaOH$ (aq) $\longrightarrow 2H_2O_{(l)} + Na_2SO_4$ (aq)

2HOH

DR

aluminum + aqueous copper (II) nitrate → copper + aqueous aluminum nitrate

$2Al_{(s)} + 3Cu(NO_3)_2$ (aq) $\longrightarrow 3Cu_{(s)} + 2Al(NO_3)_3$ (aq)

SR

chlorine + aqueous potassium bromide → bromine + aqueous potassium chloride

* $Cl_{(g)} + KBr_{(aq)} \rightarrow Br_{(\ell)} + KCl_{(aq)}$ ✓

SR

28. Balance the following chemical reactions:

a. $1\ Ni_{(s)} + 2\ HCl_{(aq)} \rightarrow 1\ NiCl_{2(aq)} + 1\ H_{2(g)}$

b. $1\ Ca(OH)_{2(s)} + 2\ HCl_{(aq)} \rightarrow 1\ CaCl_{2(aq)} + 2\ HOH_{(g)}$

c. $1\ Cl_{2(s)} + 2\ NaBr_{(aq)} \rightarrow 1\ Br_{2(l)} + 2\ NaCl_{(aq)}$

d. $2\ Cr_2O_{3(s)} \rightarrow 4\ Cr_{(s)} + 3\ O_{2(g)}$

e. $2\ Fe_{(s)} + 6\ HCl_{(aq)} \rightarrow 2\ FeCl_{3(aq)} + 3\ H_{2(g)}$

C - 6 / H - 12 / Cl - 18 | C - 6 / H - 12 / Cl - 18

(f.) $2\ C_3H_{6(g)} + 9\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6\ H_2O_{(g)}$

g. $1\ P_{4(s)} + 6\ F_{2(g)} \rightarrow 4\ PF_{3(l)}$

29. For each of the following, **classify** the reaction type and predict the **balanced chemical equation**. Provide the **word equation** as well.

a. $4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_3$

(formation)

aluminum (s) + oxygen (g) → aluminum oxide

b. $2Ag_2O_{(s)} \rightarrow 4Ag_{(s)} + O_{2(g)}$

(decomposition)

Silver oxide (s) → Silver (s) + oxygen (g)

(c.) $Br_{2(l)} + 2KI_{(aq)} \rightarrow I_{2(s)} + 2KBr_{(aq)}$

(single replacement)

bromine + potassium iodide → iodine + potassium bromide

d. A strip of zinc metal is placed into a copper (II) nitrate solution.

$Zn_{(s)} + Cu(NO_3)_{2\,(aq)} \longrightarrow Cu_{(s)} + Zn(NO_3)_{2\,(aq)}$

Zinc + copper (II) nitrate ⟶ copper (II) + zinc nitrate

- Single replacement

e. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow 2NaCl_{(aq)} + BaSO_{4\,(s)}$

barium chloride + sodium sulfate ⟶ sodium chloride + barium sulfate

- double replacement.

23. Calculate the molar mass of the following substances *

Na_2SO_4 $2 \times Na = 2 \times 22.99 = 45.98$ $1 \times S = 1 \times 32.07 = 32.07$ $4 \times O = 4 \times 16.00 = 64$ $45.98 + 32.07 + 64 = 142.05$ g/mol	H_2O $2 \times H = 2 \times 1.01 = 2.02$ $1 \times O = 1 \times 16.00 = 16.00$ $16.00 + 2.02 = 18.02$ g/mol	calcium hydroxide = $Ca(OH)_2$ $Ca \times 1 = 40.08 \times 1 = 40.08$ $O \times 2 = 2 \times 16.00 = 32$ $H \times 2 = 2 \times 1.01 = 2.02$ $40.08 + 32 + 2.02 = 74.1$ g/mol
tetraphosphorus pentaoxide ~ P_4O_5 $P \times 4 = 4 \times 30.97 = 123.88$ $O \times 5 = 5 \times 16.00 = 80$ $123.88 + 80 = 203.88$ g/mol	$MgSiO_3$ $1 \times Mg = 1 \times 24.31 = 24.31$ $1 \times Si = 1 \times 28.09 = 28.09$ $O \times 3 = 3 \times 16.00 = 48$ $24.31 + 28.09 + 48 = 100.4$ g/mol	NH_3 ammonia $1 \times N = 1 \times 14.01 = 14.01$ $H \times 3 = 3 \times 1.01 = 3.03$ $14.01 + 3.03 = 17.04$ g/mol

24. Answer the questions in the following table.

How many moles in 15 g of silicon? $n = \frac{m}{M}$, so $n = \frac{15}{28.09}$ $m = 15$g $M = 28.09$ mol $n = 0.533997861...$ $n = 0.53$ mols	How many moles in 1670 g of hydrocholic acid? $n = \frac{m}{M}$, so $n = \frac{1670}{36.46}$ $m = 1670$g $M = 1.01 + 35.45 = 36.46$ $n = 45.80362041...$ $n = 45.80$ moles	How many moles in 25 g of sodium chloride? $n = \frac{m}{M}$, so $n = \frac{25}{58.44}$ $m = 25$g $M = 22.99 + 35.45 = 58.44$ $n = 0.427789185...$ $n = 0.43$ mols

What is the mass in grams of 7.30 mol of calcium chloride?	What is the mass in grams of 5.32 mol of calcium hydroxide?	What is the mass in grams of 3.00 mol of carbon monoxide?
$m = nM$ $CaCl_2$ $n = 7.30$ $M = (2 \times 35.45) + 40.08$ $= 110.98$ $7.30 \times 110.98 = 810.154$ $= 810\ g$	$Ca(OH)_2$ $m = nM$ $n = 5.32$ $M = 40.08 + (2 \times 1.01) + (16 \times 2)$ $= 74.1\ g/mol$ $m\ 74.1\ g/1 \times 5.32 = 394.212$ $m = 394\ g$	CO $n = 3.00\ mol$ $M = 12.01 + 16 = 28.01$ $m = Mn$ $28.01 \times 3 = 84.03$ $m = 84.03\ g$

Switch the 2 metals

Unit 2: Physics

1. Define work. Include the work equation.

   The transfer of mechanical energy from one object to another.

   ~ $W = f \times d$, or work = force x distance.

2. A weightlifter exerts a force of 883 N on a barbell over a distance of 0.65 m. How much work did the weightlifter do on the barbell?

   $W = f \times d$

   $f = 883\ N$

   $d = 0.65\ m$

   $W = 883 \times 0.65 = 573.95$

   $W = 5.7 \times 10^2\ J$

3. With a single pulley, you lift a crate. If you exerted a force of 455 N and did 3276 J of work, how far did you lift the crate?

   W / f d

   $W = f \times d$, so $d = \frac{W}{f}$, so $d = \frac{3276\ J}{455\ N}$, $= 7.2\ m$

   $W = 3276\ J$

   $f = 455\ N$

   $d = 7.20\ m$

4. What are the two laws of thermodynamics? Why can't perpetual motion machines exist?

   #1 ~ Energy cannot be created or destroyed, but can be transferred from one form to another or transferred from one thing to another.

   #2 ~ Cannot convert 100% thermal energy to work

   ~