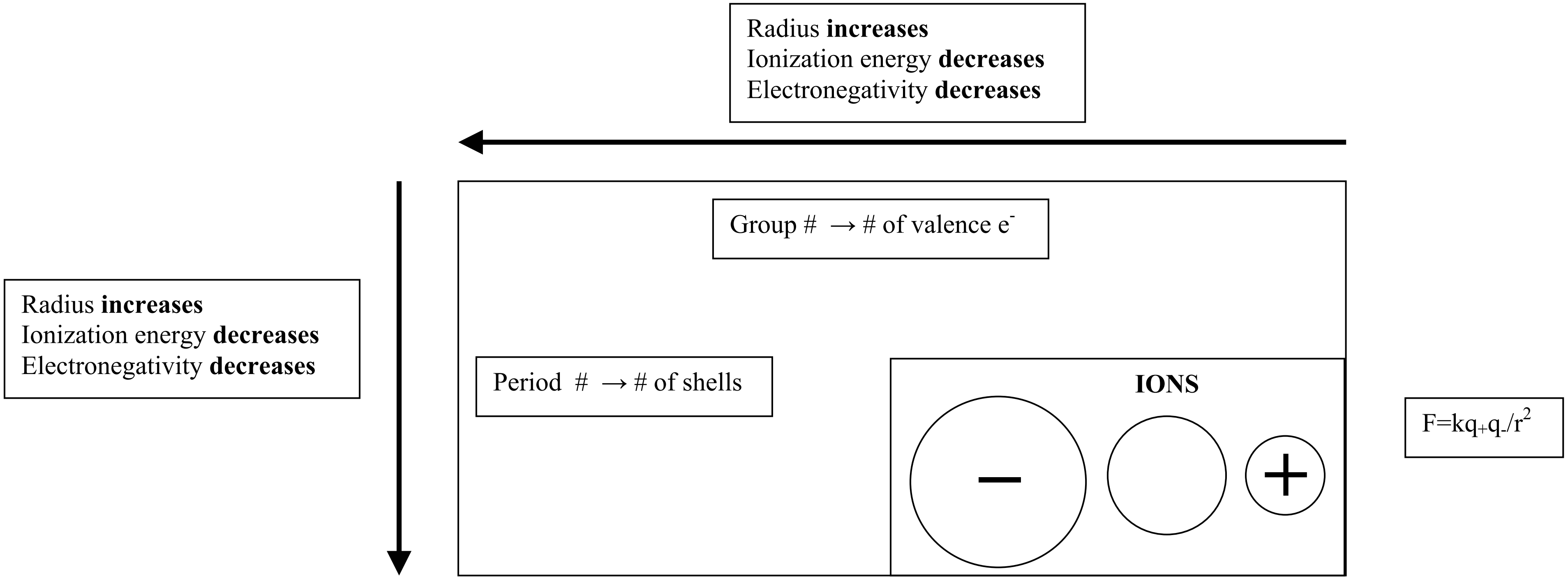
HANDOUT

Module:	Chemistry SL/HL	Topic:	**Properties of the periodic table**
Tutor:	Dr. Liakatas	Date given:	



HANDOUT

EXPLANATION OF THE PROPERTIES

In same GROUP **going down**:
► more shells → valence e^- farther from nucleus → larger atomic radius
► more shells → valence e^- farther from nucleus → weaker attraction from nucleus → smaller energy to remove an e^- → smaller ioniz. energy
► more shells → valence e^- farther from nucleus → weaker attraction of an e^- in a covalent bond → smaller electronegativity

In same PERIOD **going left**:
► valence e^- in same energy level but less protons in nucleus → weaker attraction of each valence e^- by nucleus → larger atomic radius
► valence e^- in same energy level but less protons in nucleus → weaker attraction of each valence e^- by nucleus → smaller energy to remove an e^- → smaller ioniz. energy
► valence e^- in same energy level but less protons in nucleus → weaker attraction of each valence e^- by nucleus → larger atomic radius → weaker attraction of an e^- in a covalent bond → smaller electronegativity

Negative ion: ► gain of valence e^- but same protons in nucleus → same attraction of each valence e^- by nucleus but larger e-e repulsion → larger ionic radius
Positive ion: either ► loss of all valence e^- → one shell less → smaller ionic radius
Or ► loss of some valence e^- but same protons in nucleus → same attraction of each valence e^- by nucleus but smaller e-e repulsion → smaller ionic radius

HANDOUT

Period 3:
Electrical conductivity **increases and then decreases …[5]**
Melting point/b.p. **increases and then decreases…[6]**

Group # → # of valence e^-

Group 1:
Reactivity **increases…[1]**
Melting point/b.p. **decreases…[2]**

Period # → # of shells

Group 7:
Reactivity **increases…[3]**
Melting point/b.p. **decreases…[4]**

[1]… because valence e^- farther from nucleus → easier to be lost to form cation → easier to react with an anion → larger reactivity
[2]… because bigger nucleus but same valence e^- forming free electron cloud trying to keep cations together → weaker forces → smaller m.p./b.p.
[3]… because larger electronegativity → easier to attract electron and become anion → easier to react with a cation → larger reactivity
[4]… because lower molar mass → weaker Van der Waals forces → smaller m.p./b.p.
[5]… because for metals, less electrons in valence shell to form free electron cloud **but then** for non-metals, no free electrons at all
[6]… because for metals, more valence e^- forming free electron cloud keeping cations together **but then** for non-metals, covalent bond weaker than metallic

Chemical Properties of Period 3

Chlorides of Period 3

ionic compounds	Na	NaCl	+ H_2O	→	Na^+	+ Cl^-	dissociation (conductive when molten)
	Mg	$MgCl_2$	+ H_2O	→	Mg^{2+}	+ 2 Cl^-	
covalent compounds	Al	Al_2Cl_6 ($AlCl_3$)	+ 3 H_2O	→	Al_2O_3 (oxide)	+ 6 HCl	acidic solutions (not conductive)
	Si	$SiCl_4$	+ 4 H_2O	→	$Si(OH)_4$ (base)	+ 4 HCl	
	P	PCl_3	+ 3 H_2O	→	H_3PO_4 (acid)	+ 3 HCl	
	S	-					
	Cl	Cl_2	+ H_2O	↔	HClO	+ HCl	

Oxides of Period 3

ionic compounds	Na	Na_2O	+ H_2O	→	2 Na^+	+ 2 OH^-	form bases
	Mg	MgO	+ H_2O	→	$Mg(OH)_{2(s)}$		
giant covalent compounds	Al	Al_2O_3 Al_2O_3	+ 6 HCl + 2 NaOH	→	2 $AlCl_3$ 2 $NaAl(OH)_4$	+ 3 H_2O (base) + 3 H_2O (acid)	is amphoteric
	Si	SiO_2	+ 2 NaOH	→	Na_2SiO_3	+ H_2O	is weak acid
covalent compounds	P	P_4O_{10}	+ 6 H_2O	→	4 H_3PO_4		form acids
	S	SO_3	+ H_2O	→	H_2SO_4		
	Cl	Cl_2O_7	+ H_2O	→	2 $HClO_4$		

Extra ½ O for each successive oxide

HANDOUT

Properties of d-block elements

1. **Multiple oxidation numbers** because the 3d and 4s orbitals have similar energy.

2. **Formation of complex ions** (mainly Co, Ni, Cu) when a ligand (usually H_2O, NH_3, Cl) forms a dative (coordinate) bond
   H_2O, NH_3 → octahedral, Cl → tetrahedral

3. **Colored** because the d orbitals split and electrons can absorb light and rise to a higher energy orbital
   Sc & Zn ions are colorless

4. **Catalysts**:

Decomposition of hydrogen peroxide	$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$	(catalyst = MnO_2)
Hydrogenation of alkenes	$C_2H_4 + H_2 \rightarrow C_2H_6$	(catalyst = Ni)
Haber process	$N_2 + 3\ H_2 \leftrightarrow 2\ NH_3$	(catalyst = Fe)
Contact process	$2\ SO_2 + O_2 \rightarrow 2\ SO_3$	(catalyst = V_2O_5)

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	F	Cl	Br	I
State at room temperature	Gas	Gas	Liquid	Solid
Color of ions	Colorless	Colorless	Colorless	Colorless
Color at normal state	Pale yellow	Yellow/green	Red/brown	Black/purple
Color in solution		Green	Yellow/orange/brown (depending on concentration)	Brown
+ Ag	X	AgCl → **white** precipitate (turns black in sunlight)	AgBr → **creamy** precipitate	AgI → **yellow** precipitate
+ Cl_2	X	X	$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$	$Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$
+ Br_2	X	X	X	$Br_2 + 2I^- \rightarrow I_2 + 2Br^-$

Intermolecular forces

Strength increases ↓

Van der Waals:
- ► attraction of temporary dipoles due to random motion of valence electrons
- ► in all species, polar and non-polar
- ► proportional to molar mass and molecule's surface
- ► explains: forces between halogen molecules, increasing m.p. of halogens, increased m.p. of straight chain isomers

Dipole-dipole:
- ► electrostatic attraction between polar (asymmetric) molecules
- ► explains: increased m.p. of hydrogen halides (HX) compared to noble gases

Hydrogen bond:
- ► between a H attached to a F, O, N and a free electron pair of a F, O, N
- ► explains: high m.p. of water, higher m.p. of alcohols and organic acids