

ADVANCED CHEMISTRY

INORGANIC CHEMISTRY HW #① - WINTER

Chapter 2

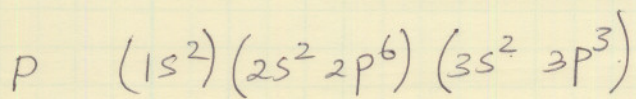
- ⑩ (a) $5d$ $l=2$ $m_l = -2, -1, 0, +1, +2$
 $4f$ $l=3$ $m_l = -3, -2, -1, 0, +1, +2, +3$
 $7g$ $l=4$ $m_l = -4, -3, -2, -1, 0, +1, +2, +3, +4$

(b) $3d$ $n=3$ $l=2$ $m_l = -2, -1, 0, +1, +2$
In each case $m_s = \pm \frac{1}{2}$

- ⑫ (a) V(23) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
(b) Br(35) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
(c) Ru(44) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^6$
 Ru^{3+} $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^5$
(d) Hg(80) $[Xe] 6s^2 5d^{10} 4f^{14}$
 Hg^{2+} $[Xe] 4f^{14} 5d^{10}$
(e) Sb(51) $[Kr] 5s^2 4d^{10} 5p^3$

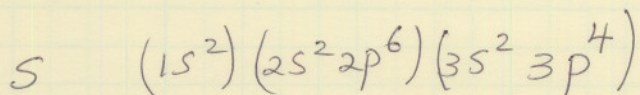
- ⑬ Mn^{2+} has five 3d electrons
 Ti^{2+} has two 3d electrons

~~14~~
15



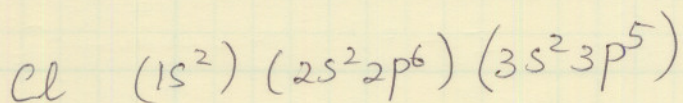
$$S = 4(0.35) + 8(0.85) + 2(1.0) = 10.2$$

$$Z^* = Z - S = 15 - 10.2 = \underline{\underline{4.8}}$$



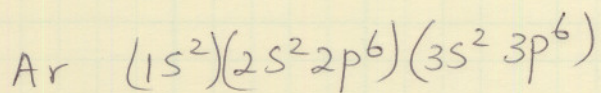
$$S = 5(0.35) + 8(0.85) + 2(1.0) = 10.55$$

$$Z^* = Z - S = 16 - 10.55 = \underline{\underline{5.45}}$$



$$S = 6(0.35) + 8(0.85) + 2(1.0) = 10.90$$

$$Z^* = Z - S = 17 - 10.90 = \underline{\underline{6.10}}$$



$$S = 7(0.35) + 8(0.85) + 2(1.0) = 11.25$$

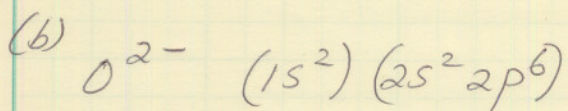
$$Z^* = Z - S = 18 - 11.25 = \underline{\underline{6.75}}$$

P S Cl Ar

—————→ atomic radius decreases

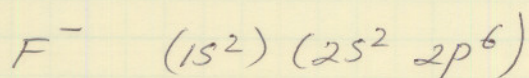
—————→ Z^* increases

since valence electrons are filling into the same orbital (3p orbital), as Z^* increases the atomic radius must decrease. Indeed this is



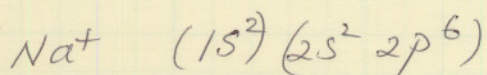
$$S = 7(0.35) + 2(0.85) = 4.15$$

$$Z^* = Z - S = 8 - 4.15 = \underline{\underline{3.85}}$$



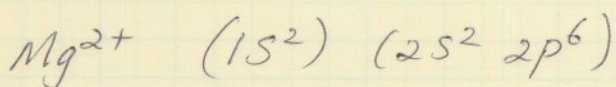
$$S = 7(0.35) + 2(0.85) = 4.15$$

$$Z^* = Z - S = 9 - 4.15 = \underline{\underline{4.85}}$$



$$S = 7(0.35) + 2(0.85) = 4.15$$

$$Z^* = Z - S = 11 - 4.15 = \underline{\underline{6.85}}$$

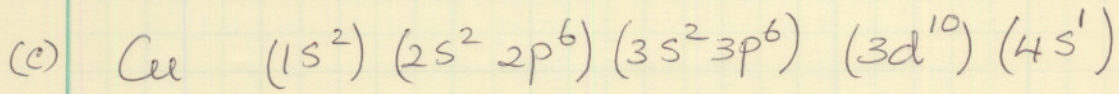


$$S = 7(0.35) + 2(0.85) = 4.15$$

$$Z^* = Z - S = 12 - 4.15 = \underline{\underline{7.85}}$$

all these ions are isoelectronic ions. Therefore as the charge of the nucleus increases, size of the ion must decrease, which means that the effective nuclear charge must increase (this is consistent with the observation).

	O^{2-}	F^-	Na^+	Mg^{2+}
# of p	6	7	8	9
# of \bar{e}	8	8	8	8
size	→ size decreases			
	→ Z^* increases.			



4s electron

$$S = 18(0.85) + 10(1.0) = 25.3$$

$$Z^* = Z - S = 29 - 25.3 = \underline{\underline{3.7}}$$

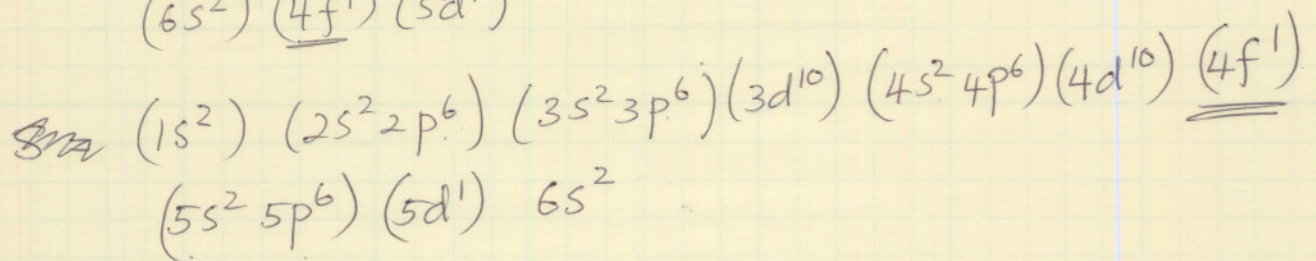
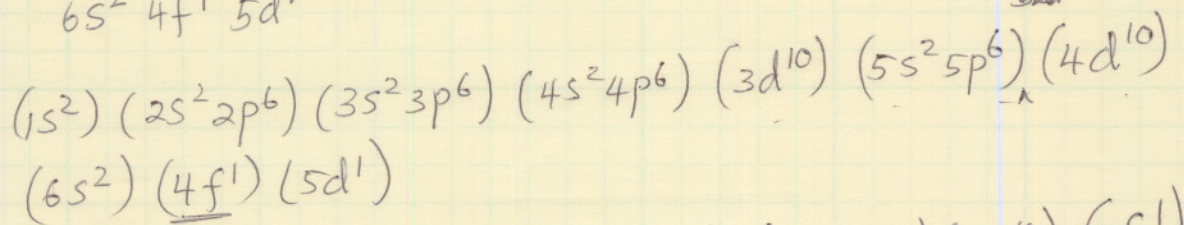
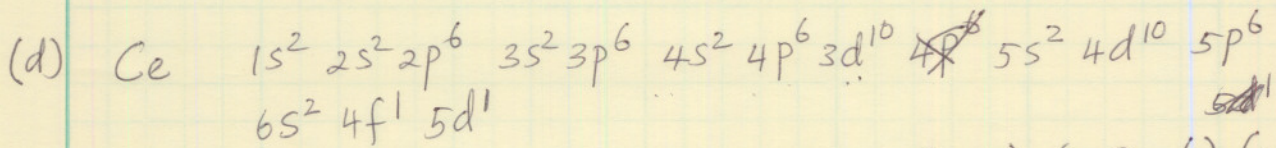
3d electron

$$S = 9(0.35) + 18(1.0) = 21.15$$

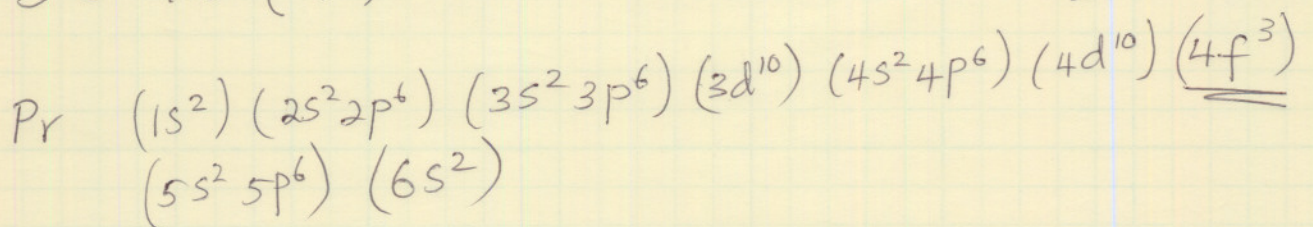
$$Z^* = Z - S = 29 - 21.15 = \underline{\underline{7.85}}$$

Z^* for the 4s electron \ll Z^* for the 3d electron

\therefore 4s electron is likely to be lost compared to the 3d electron.



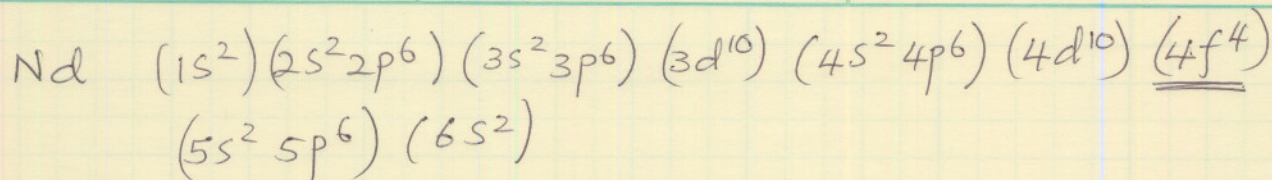
$$S = 1.00(46) = 46 \quad Z^* = 58 - 46 = \underline{\underline{12}}$$



$$S = 2(0.35) + 46(1.00) = 46.7$$

$$Z^* = 59 - 46.7$$

$$Z^* = \underline{\underline{12.3}}$$



$$S = 3(0.35) + 46(1.00) = 47.05$$

$$Z^* = Z - S = 60 - 47.05 = \underline{\underline{12.95}}$$



atomic #

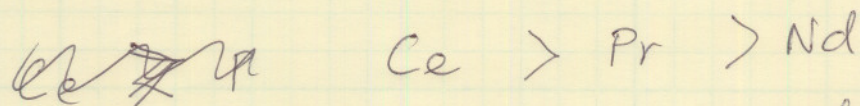
Ce Pr Nd

58 < 59 < 60

Z^*

12 < 12.3 < 12.95 ~~12.3~~

∴ atomic radius must decrease as Z^* increases



—————→ size decreases

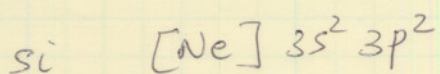
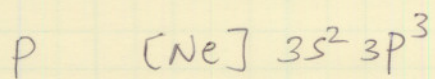
This is consistent with lanthanide contraction.

(16)

(a) Ga > Ca as you go across the same period atomic size decreases, valence e⁻ are closer to the nucleus, therefore I.E. increases.

(b) Mg > Ca as you go down a group, atomic radius increases (Ca is larger than Mg). Hence it is easier to remove an e⁻ from Ca than from Mg.

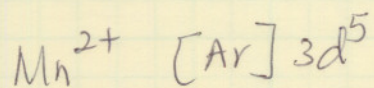
(c) $P < Si$



Since the 3p orbital is exactly half-filled in P, it is a stable configuration compared with that of Si.

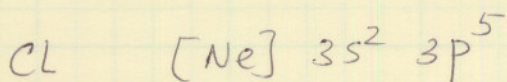
∴ P is less electronegative than Si.

(d) Mn $[Ar] 4s^2 3d^5$



The 4s electrons experience lower effective nuclear charge (Z^*) than 3d electrons, hence 4s electrons are easier to remove when forming Mn^{2+} ions.

(22) (a) S $[Ne] 3s^2 3p^4$



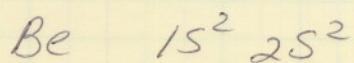
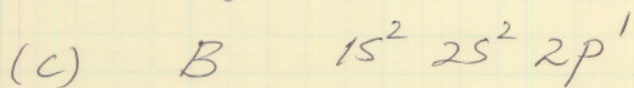
Cl needs only one more e^- to achieve the rare gas electron configuration (S needs 2 more e^- 's).

Therefore Cl has a higher e^- affinity than S.

(b) I is larger in size than Br since it is lower in Group 7A than Br.

∴ and atomic radius of I $>$ atomic radius of Br
 ∴ therefore I does not have a higher electron

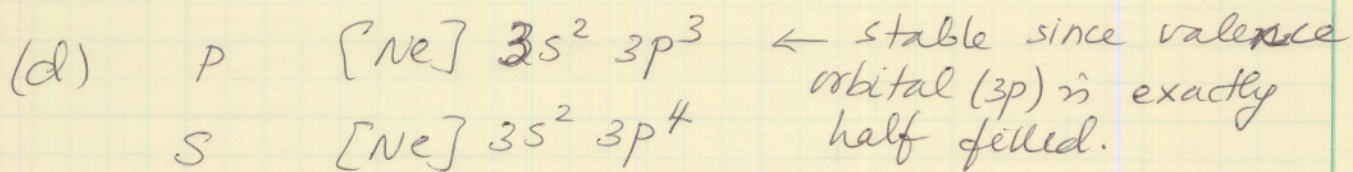
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affinity compared with Br.



Be's valence orbital (2s) is completely filled. This is a stable configuration.

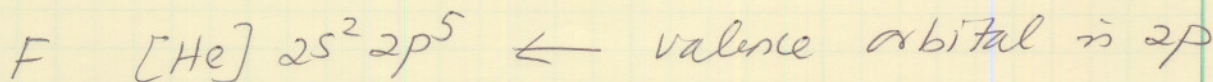
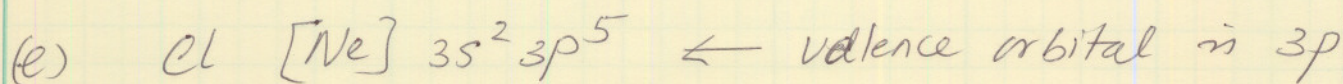
B has one e^- in its valence orbital (2p), ~~at~~ which can be given up more easily compared with the 2s orbital of Be.

\therefore B has a lower I.E. than Be.



\therefore Easier to remove an e^- from 3p of S than from 3p of P.

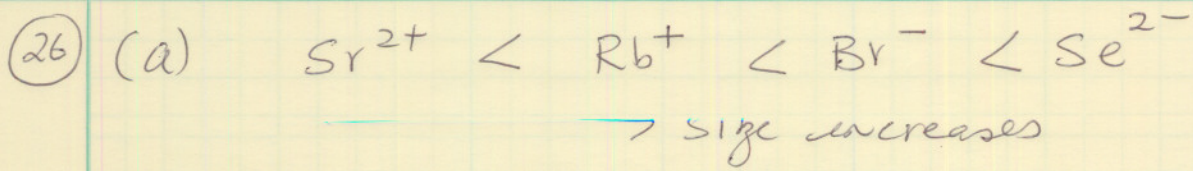
\therefore S has a lower I.E. than P.



Cl has a larger atomic radius than F.

\Rightarrow easier to remove the valence e^- from Cl compared with that of F.

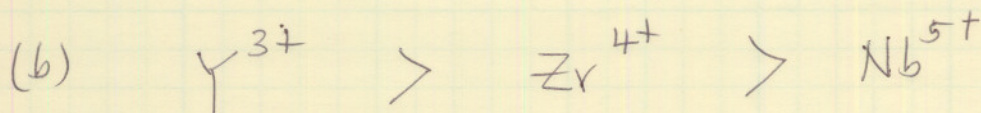
\therefore I.E. of Cl < I.E. of F



	Sr^{2+}	Rb^+	Br^-	Se^{2-}	
# of \bar{e}	36	36	36	36	← isoelectronic ions.
# of p	38	37	35	38 34	

Sr^{2+} ~~Se^{2-}~~ has the highest p: \bar{e} ratio. \Rightarrow smallest in size

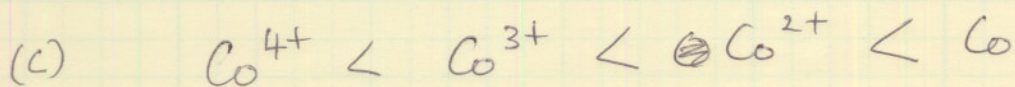
Se^{2-} has the lowest p: \bar{e} ratio \Rightarrow largest in size.



	Y^{3+}	Zr^{4+}	Nb^{5+}	
# of \bar{e}	36	36	36	isoelectronic ions.
# of p	39	40	41	

p: \bar{e} ratio is highest for $\text{Nb}^{5+} \Rightarrow$ smallest in size

p: \bar{e} ratio is lowest for $\text{Y}^{3+} \Rightarrow$ largest in size.



	Co^{4+}	Co^{3+}	Co^{2+}	Co
# of p	27	27	27	27
# of \bar{e}	23	24	25	27

highest p: \bar{e} ratio is in $\text{Co}^{4+} \Rightarrow$ smallest in size

lowest p: \bar{e} ratio is in $\text{Co} \Rightarrow$ largest in size.